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LECTURE NOTES

ON

THE METALS

STODDARD.

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OUTLINE

OF

LECTURE NOTES

ON

GENERAL CHEMISTRY

BY

JOHN T. STODDARD, PH. D.
PROFESSOR OF CHEMISTRY IN SMITH COLLEGE

THE METALS

A. S. BARNES & COMPANY
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THE METALS.

Some fifty elements are *metals.* No sharp line can be drawn between the metals and the non-metals, and the classification into these two groups is to some extent arbitrary, and is retained chiefly for convenience. The metals exhibit, however, certain properties which serve to distinguish them in a general way from the non-metals; chief among these are the physical property of metallic lustre and the chemical property of forming salts by the replacement of the hydrogen of acids.

OCCURRENCE.

Many of the metallic elements are widely distributed, but only some fifteen or sixteen occur abundantly. Very few are found in the free or *native* state. Of these, gold and the metals of the platinum group almost always occur native; bismuth is usually found in the metallic state; while silver, copper, mercury and antimony are sometimes free, but oftener combined.

The compounds in which the metals occur most commonly, are: for the heavy metals, oxides and sulphides; for the light metals, silicates, carbonates, sulphates, and other salts.

An ore is a natural metallic compound from which the metal may be profitably extracted.

EXTRACTION AND PREPARATION.

The ores are prepared by a more or less perfect separation from the rock in which they almost invariably occur. The method of extraction varies widely in different cases ; but may be indicated in a general way :—

Oxides are reduced by carbon (coal or charcoal) at a high temperature.

Sulphides are converted into oxides by roasting, and then reduced as above ; or after partial oxidation are reduced by reaction with the unchanged sulphides.

Quite a number of metals are prepared from their chlorides by electrolysis or by the reduction with sodium or potassium.

PHYSICAL PROPERTIES.

State of Aggregation. All of the metals except mercury are solid at ordinary temperatures ; all except osmium can be fused, and a number distilled. (Cf. tables in appendix.)

Opacity and Lustre. All metals are opaque except in very thin films, and exhibit when polished a *metallic lustre.*

Color. The color of nearly all of the metals is white or gray. Copper (red), gold, strontium and calcium (yellow) are the only striking exceptions.

In a finely divided state—as by reduction from their salts—the metals form dull grey or black powders which, however, recover their metallic lustre and color when rubbed or fused.

Odor and Taste. A few metals yield a peculiar odor when rubbed, and when tarnished have an acrid "metallic" taste.

Crystallization. Most metals can be crystallized. The crystals are obtained from the molten metal or its vapor on cooling; or from a solution of one of its salts by electrolysis or reduction by another metal. The crystals are usually forms belonging to the regular system, such as cubes or octahedrons. The *fracture of metals* is described as:—crystalline, granular, fibrous, silky, columnar, or conchoidal.

Hardness, Tenacity, Malleability and Ductility. In these respects the metals differ very widely from each other. (Cf. tables in appendix.)

Conduction of Heat and Electricity. The metals are far better conductors of both heat and electricity than most of the non-metals or any chemical compounds. The conductivity varies considerably in different metals, and also in the same metal according to its physical state and purity. (Cf. appendix.)

Specific gravity of the metals varies widely. (Cf. appendix.) Those having a specific gravity above 5 are called *heavy metals,* the rest *light metals.*

Solubility. The metals are insoluble as such in water and ordinary solvents. They dissolve, however, in each other when melted, forming alloys.

Specific and Atomic Heat. The product of the specific heats of the metals by the corresponding atomic weights is an approximately constant number (6.4), which expresses the capacity for heat of the atoms; thus *all metals have the same atomic heat.* Hence the determination of the specific heat is a means of checking the atomic weight of a metal, or of ascertaining it in a doubtful case. Certain non-metals have the same atomic heat as the metals, viz.: chlorine, bromine, iodine, selenium, tellurium and arsenic.

CHEMICAL PROPERTIES.

The metals differ widely in their chemical as well as in their physical properties, but are all *electro-positive* with reference to the non-metals (Cf. appendix); form *basic* oxides, and salts by the replacement of the hydrogen of acids. Other generic chemical properties are given under Compounds of the Metals.

HISTORY.

The metals gold, silver, copper, iron, tin, lead and mercury were known to the ancients. All but four — antimony, cobalt, platinum and zinc — of the others, have been discovered since the middle of the last century.

COMPOUNDS OF THE METALS.

Alloys. Most metals unite with each other on fusion, forming *alloys,* which appear in some cases to be definite chemical compounds, but are usually to be regarded as a mixture of such compounds with an excess of one of the metals.

Alloys are always metallic in character; they are all (except one of potassium and sodium) solid at ordinary temperatures; usually melt lower than their constituent metals and often differ widely from these in their hardness, malleability, ductility, specific gravity and behavior toward chemical influences. These properties are in general not the mean of those of the constituent metals, but different and peculiar to the alloy, and are greatly influenced by the proportions employed. Alloys are of the greatest practical value, furnishing an indefinite variety of "metals" adapted to special purposes, for which no single metal would serve. (Cf. appendix.)

Amalgams are alloys of metals with mercury. (Cf. appendix.)

OXIDES AND HYDROXIDES.

All the metals combine with oxygen, many of them uniting with it in more than one proportion. Some of the metallic oxides occur naturally as crystalline minerals, and in two or three cases are the most important ores of the metals they contain. Most of the oxides, however, are obtained only by artificial means.

Preparation. The more important methods of preparing the oxides of the metals are :—

a. Direct oxidation of the metal in air or oxygen, usually at a more or less elevated temperature. All the metals except gold, silver and platinum can be thus oxidized. By the combined action of moisture and air many metals, which are unchanged in dry air at ordinary temperatures, are converted into oxides or hydroxides (e. g. rusting of iron). In most cases this change is merely superficial (e. g. lead).

b. Action of oxidizing agents, e. g. HNO_3, KNO_3.

c. Most metals are converted into oxides or hydroxides by the action of H_2O. In some cases this takes place at ordinary temperatures (e. g. alkali metals); in others only by the aid of heat.

d. Oxides or hydroxides are precipitated from metallic solutions by the alkaline hydroxides.

e. All hydroxides except those of the alkalies and barium are converted by heat into oxides.

f. Oxides are formed by the decomposition of certain salts, especially carbonates and nitrates, by heat.

Properties. The artificially prepared oxides and hydroxides are mostly amorphous, lusterless powders or masses, often characteristically colored; of less specific gravity and less fusible than the metals from which they are derived; non-conductors of electricity. All, except those of the alkalies and alkaline earths, $TlOH$ and CrO_3, are insoluble in water, but dissolve in acids. A few are completely decomposed by heat into metal and oxygen; others lose part of their oxygen and are changed to a lower oxide; and others still are unaltered. Many are reduced by hydrogen, and nearly all by carbon or the alkali metals.

Oxides are classed as:—

Basic oxides, which when treated with acids, form salts and water:—e. g. $CaO+2HCl=H_2O+CaCl_2$.

The soluble oxides and hydroxides of this class have an alkaline reaction. Many metals yield two or more basic oxides and corresponding hydroxides.

Peroxides, which contain more oxygen than the basic oxides and give with hydrochloric acid either chlorine or hydrogen peroxide :—e. g.

a. $MnO_2+4HCl=MnCl_2+2H_2O+Cl_2$,

b. $BaO_2+2HCl=BaCl_2+H_2O_2$,

and with sulphuric acid either oxygen or hydrogen peroxide.

Acid-forming oxides containing more oxygen than the peroxides and uniting with other metallic oxides to form salt-like compounds.

The same oxide sometimes acts as basic, acid-forming or peroxide in different cases.

SULPHIDES.

Sulphides of many of the heavy metals occur naturally and in several instances constitute the most important sources of the metals. Sulphides of all of the metals can be formed artificially, and in the case of the metals of the alkalies and alkaline earths *hydrosulphides* can be prepared.

Preparation. a. By direct union of the metal and sulphur, usually by the aid of heat. In this way almost all the metallic sulphides can be formed.

b. By heating the oxide in hydrogen sulphide.

c. By precipitating a solution of a metallic salt by hydrogen sulphide or an alkaline hydrosulphide. Many sulphides cannot be formed in this wet way.

d. By heating the sulphate with carbon.

Properties. The sulphides are all solids, many of them characteristically colored. They conduct electricity; are fusible—those of the light metals less and those of the heavy metals more so than the metals from which they are derived. Heated in the air, most sulphides are oxidized, being converted either into oxides or sulphates (roasting). A few can be reduced by heat alone, many by hydrogen or metals at an elevated temperature, but not by carbon.

SALTS.

A salt is a compound formed by the replacement of the whole or part of the hydrogen of an acid by a metal.

Normal salts are those in which all of the replaceable hydrogen is replaced by a metal. E. g. K_2SO_4.

Acid salts are those which still contain replaceable hydrogen. E. g. $KHSO_4$.

Basic Salts are normal salts in which one or more acid groups have been replaced by hydroxyl (HO) or oxygen. They are thus compounds which present the characteristics of both salts and bases. E. g. $HOPbNO_3$, $HOZnCl$.

Double salts are such as contain two or more different metals. They may be normal or acid. E. g. $MgNH_4PO_4$, $NaNH_4HPO_4$, $AlK2SO_4$.

Salts are prepared in general :—a. By the action of the acid upon the metal, its oxide, hydroxide or carbonate. b. By metathetical reaction of two salts, usually in solution, and separation of the new salts formed by means of their difference in solubility.

Salts are almost without exception solid substances. Many dissolve in water, as a rule more freely in hot than in cold water, and hence separate out in crystalline form when a hot saturated solution is allowed to cool. (Cf. appendix, table of solubilities.)

THE METALS OF THE ALKALIES.

The metals of this group (Potassium, Sodium, Lithium, Rubidium, Caesium,) are all strongly positive monads, which decompose water energetically at all temperatures, with the formation of *alkaline* hydroxides. Their compounds are with few exceptions very soluble in water. The ammonium salts resemble the salts of the alkalies so closely that they are placed in this group.

POTASSIUM. K^{I} (Kalium)=39.1

Occurs only in its salts. Chiefly as silicate in felspar and many other minerals. By the gradual weathering and disintegration of these, its salts have become widely distributed in the soil, are taken up by plants and form the principal soluble part of their ashes. Potassium occurs also as chloride in considerable deposits and as nitrate (saltpetre).

Formed: a. From its hydroxide, chloride or cyanide by electrolysis. b. From its hydroxide by the action of iron at a white heat. c. Usually *prepared* by heating an intimate mixture of potassium carbonate and carbon (obtained by calcining crude tartar) to a white heat in an iron retort :—

$$K_2CO_3+2C=3CO+K_2.$$

The potassium vapor is cooled rapidly in a flattened iron condenser to avoid the formation of an explosive compound ($K_2C_2O_2$) which is otherwise produced, and drops in the liquid state into a vessel of petroleum, where it solidifies.

Properties. Silver-white lustrous metal. Specific gravity 0.87. Brittle at 0°, at 15° soft as wax and can be welded, melts at 62°.5 and volatilizes at a red heat, yielding a green vapor. Unaltered in dry pure air, but tarnishes quickly in ordinary air with the formation of hydroxide and carbonate. Heated, it burns with a violet flame. Fuses and takes fire when thrown on water, decomposing the latter with formation of potassium hydroxide and hydrogen. Is a powerful reducing agent, withdrawing oxygen, chlorine, etc., from most of their compounds.

Uses. Potassium has been used on account of its powerful reducing power for the preparation of certain elements from their compounds, but is now replaced for this purpose by the cheaper sodium. Employed in the preparation of pure potassium hydroxide.

History. Discovered by Davy, 1807, by electrolysis of potassium hydroxide. Soon after obtained by Gay Lussac and Thenard by chemical means.

COMPOUNDS OF POTASSIUM.

Potassium oxides. *Potassium monoxide,* K_2O, and *peroxide,* K_2O_4, are both formed when the metal burns in dry air. K_2O is also obtained by heating the metal with potassium hydroxide in an atmosphere of hydrogen ; K_2O_4 by heating the metal first in dry air and then in oxygen. They are both deliquescent solids, which unite energetically with water forming the hydroxide.

Potassium Hydroxide, KOH. *Prepared*: a. Pure, by dissolving the metal or its oxide in water. b. Generally, by boiling a solution of potassium carbonate with slaked lime :— $K_2CO_3 + Ca(OH)_2 = CaCO_3 + 2KOH$. The clear solution is evaporated to dryness, and the hydroxide fused and cast in sticks.

Properties. White brittle solid, which melts below a red heat to an oily liquid. Very soluble in water and alcohol. In the air it deliquesces and is converted into carbonate by the absorption of carbon dioxide. It is a powerful base, neutralizing the strongest acids and displacing most other bases from their salts. Its concentrated solution attacks and destroys organic tissues and corrodes glass.

Uses. Potassium hydroxide or caustic potash is used in making soft soap, in the preparation of oxalic acid from saw-dust, and for some minor purposes. It is an important reagent in the laboratory and is used as a cautery in surgery.

Potassium Chloride, KCl. *Occurs* in large quantities as sylvine (KCl), and in carnallite ($KCl, MgCl_2 + 6H_2O$) in the Stassfurt salt deposits ; is present in sea-water and saline springs.

Formed by the direct union of its elements and in many chemical reactions, but usually obtained from the natural sources.

Properties and uses. Colorless cubes soluble in water; melts at a red heat and volatilizes at a slightly higher temperature.

Potassium chloride is the starting point for the preparation of most of the other potassium compounds. It is employed in an impure state as a fertilizer.

Potassium Bromide, KBr. Occurs naturally in small quantities in sea-water and certain mineral springs. Prepared by the action of bromine on caustic potash:

$$6KOH+3Br_2=5KBr+KBrO_3+3H_2O.$$

If this mixture is evaporated and gently ignited, the bromate is converted into bromide.

Potassium bromide resembles the chloride. It is used in photography and is a valuable medicine.

Potassium Iodide, KI. Occurs in sea-water and saline springs, and is a constituent of the ash of marine plants. Prepared as the bromide, and resembles this and the chloride. Used largely in photography and medicine.

Potassium Fluoride, KF, is formed by neutralizing aqueous hydrofluoric acid with potassium hydroxide or carbonate. Crystallizes in cubes; is very deliquescent; its solution attacks glass; unites with other fluorides to form double salts.

Potassium Chlorate, $KClO_3$. *Formed* and formerly prepared by leading chlorine into a hot solution of caustic potash:—

$$6KOH+3Cl_2=5KCl+3H_2O+KClO_3.$$

Prepared usually from potassium chloride by reaction with calcium chlorate, which is made by saturating hot milk of lime with chlorine:—

$$Ca2ClO_3+2KCl=CaCl_2+2KClO_3.$$

The potassium chlorate separates out on account of its small solubility and is purified by recrystallization.

Properties. Large transparent monoclinic crystals. Melts at 330° and at 350° begins to decompose into KCl and oxygen. Dissolves sparingly in cold water, more readily in hot (6:100 at 15°; 60:100 at 100°). Is a powerful oxidizing agent, uniting explosively with combustible substances (e. g. sulphur, sugar, charcoal) when rubbed or heated with them.

Used for making oxygen; as an oxidizing agent in calico printing, in the mixtures employed for matches, fireworks, etc., in the laboratory; and in medicine.

Potassium Perchlorate, $KClO_4$, is prepared by heating the fused chlorate until it begins to solidify :—$2KClO_3 = KCl + O_2 + KClO_4$. It is less explosive than the chlorate and is employed to some extent instead of this in pyrotechny.

Potassium Hypochlorite, KClO, is known only in solution. By leading chlorine into a cold dilute solution of caustic potash or potassium carbonate, a mixture of KClO and KCl is obtained, which was formerly largely used for bleaching purposes ("Eau de Javelle").

Potassium Sulphates. *Normal Potassium Sulphate,* K_2SO_4, occurs in kainite, polyhalite and natural alum. It is *prepared :* a. From potassium chloride by the action of sulphuric acid. b. From potassium chloride by reaction with kieserit ($MgSO_4$) and crystallization. c. From kainite. d. It is also obtained as bye-product in the preparation of potassium bichromate and in other chemical manufactures.

Properties. Potassium sulphate crystallizes in small hard rhombic pyramids; has a bitter saline taste, is sparingly soluble in water and fuses at a high heat without decomposition.

Used in large quantities for the preparation of alum, glass, and potassium carbonate (by Leblanc's process); as a component of artificial fertilizers, and to a small extent in medicine.

Acid Potassium Sulphate, $KHSO_4$, is prepared from potassium nitrate by the action of sulphuric acid. It forms large crystals, readily soluble in water; melts at 197°; at a higher temperature loses water and is converted into *potassium pyrosulphate* ($K_2S_2O_7$) which at 600° is decomposed into K_2SO_4 and SO_3.

Potassium Nitrate, KNO_3. *Occurs* with other nitrates as an efflorescence on the soil in various hot and dry countries (India), and in caverns (Kentucky), being the product of the decomposition of nitrogenous organic matter in contact with an alkali. The natural process is imitated in saltpetre plantations by exposing animal matter mixed with wood ashes and lime to the action of the air.

Obtained from both these sources by boiling out the soil or deposit with water, adding potassium carbonate to decompose the calcium nitrate present and allowing the nitre to crystallize out.

Prepared in large quantities by boiling together solutions of sodium nitrate and potassium chloride:

$$NaNO_3+KCl=NaCl+KNO_3.$$

The sodium chloride being much less soluble than the potassium nitrate is deposited first and the latter is obtained from the clear solution by evaporation.

Properties. Crystallizes in rhombic prisms. Has a cooling saline taste. Dissolves in water with absorption of much heat. Is very soluble, particularly in hot water (1 : 32 at 20° ; 1 : 0.4 at 100°). Melts at 340° and decomposes above this temperature, yielding first potassium nitrite (KNO_2). Is a very powerful oxidizing agent.

Uses. Potassium nitrate, "nitre" or "saltpetre," is used chiefly in the manufacture of gunpowder; also in glass making, pyrotechny, salting meat, the preparation of freezing mixtures, in medicine and in many chemical operations.

History. Called saltpetre (sal-petrae) because obtained by lixiviation of certain porous rocks in India. Was formerly all obtained from natural sources. Saltpetre is made in "plantations" especially in Sweden, where every landed proprietor has to pay a tax in saltpetre, and in Switzerland.

Gunpowder consists of an intimate mixture of nitre, charcoal and sulphur. The nitre cannot be replaced by the cheaper sodium nitrate, because this is hygroscopic. The proportions of the constituents vary somewhat in different countries and according to the purpose for which the powder is made. The composition of an average gunpowder is 75 parts saltpetre, 11.5 sulphur and 13.5 charcoal.

The force of the explosion is due to the sudden evolution of large quantities of gases, whose normal volume is greatly increased by the heat of the explosion. The gases produced by the firing of gunpowder are chiefly nitrogen and carbon dioxide; the temperature is about 2200°; the *normal volume* of the gases is about 280 times that of the powder; the pressure, when the powder entirely fills the space in which it is fired, is about 42 tons per square inch. The solid products of combustion are chiefly potassium carbonate, sulphate, thiosulphate and sulphide. The decomposition of an average gunpowder fired in a closed space cannot be represented even by a comparatively complicated chemical equation.

Potassium Nitrite, KNO_2, is formed by heating the nitrate to a moderate red heat. This decomposition takes place more readily when lead is added :—

$$KNO_3+Pb=PbO+KNO_2.$$

White mass; fusible, deliquescent and very soluble. Used in making certain (azo) aniline dyes and in the laboratory for the separation of cobalt and nickel, and in organic chemistry.

Potassium Carbonate, K_2CO_3. This salt under the name of potash, or pearlash, was formerly obtained entirely (a) from wood ashes by lixiviation. It is now *prepared* not only from this source, but also in large quantities (b) from potassium chloride by a process identical with the Leblanc soda process (cf. soda); (c) from the solution formed by washing wool (suint); and (d) from the molasses from beet-root. Pure potassium carbonate was formerly obtained from pure cream of tartar (hydrogen potassium tartrate) by calcination and lixiviation; it is now prepared by heating the acid carbonate, which can be easily obtained pure.

Properties. Potassium carbonate is a white powder or mass which is deliquescent and very soluble in water. Melts at a red heat. It has a strong alkaline reaction and an alkaline taste.

Used largely in making soft-soap, crystal glass and the preparation of other potassium compounds. Its place has been taken in many industries by the much cheaper soda.

Acid Potassium Carbonate, or *bicarbonate of potash,* $KHCO_3$, is prepared by passing carbon dioxide through a concentrated solution of the normal carbonate; the acid carbonate being much less soluble in water than the normal salt, crystallizes out. Large transparent monoclinic crystals. Its solution loses carbon dioxide on boiling and the dry salt is decomposed by heat into the normal carbonate, water and carbon dioxide. It was formerly much used, under the name of *saleratus,* in baking powders and in effervescing drinks.

Potassium Sulphides. *Potassium hydrosulphide,* KSH, is obtained by saturating caustic potash solution with H_2S:—$KOH+H_2S=H_2O+KSH$. On adding to this solution an equal quantity of caustic potash, potassium monosulphide, K_2S, is formed:—$KSH+KOH=H_2O+K_2S$. Both are strong sulphur bases, combining with the sulphides of arsenic, antimony and tin to form soluble sulphur salts. Other sulphides are known, among them an impure polysulphide—*liver of sulphur*—which is used in medicine.

Potassium Cyanide, KCN, is prepared by fusing together dry potassium ferrocyanide and potassium carbonate :—$K_4Fe(CN)_6 + K_2CO_3 = Fe + KOCN + 5KCN + CO_2$. Potassium cyanide is very soluble in water ; decomposed by acids even by the carbon dioxide of the air with evolution of prussic acid (HCN). Hence the salt always smells of prussic acid and is exceedingly poisonous. It reduces many oxides when fused with them, being converted into potassium cyanate (KOCN) ; fused with sulphur or certain metallic sulphides, it is transformed into potassium thiocyanate (KSCN).

Used largely in photography and electro-plating, and as a reducing agent in the laboratory.

DETECTION AND DETERMINATION OF POTASSIUM.

Potassium is detected by the violet color which its compounds impart to a non-luminous flame. Separated from sodium by precipitation of the platinum double chloride (K_2PtCl_6) and usually determined in this form.

SODIUM. Na^I (Natrium)=23.

Occurs very abundantly and widely distributed as chloride in immense deposits (rock salt) and dissolved in sea-water, salt lakes and springs and in many mineral waters. The sulphate, nitrate, borate, carbonate, bromide and iodide also occur naturally ; the fluoride is combined with AlF_3 in cryolite, and the silicate is a constituent of many minerals. The salts of sodium are absorbed by marine plants and are essential to animal life.

Prepared from the carbonate by the action of carbon at a white heat. The sodium distils off from the mixture, is purified by redistillation, and is then melted and cast into sticks.

Properties. Sodium resembles potassium. Specific gravity 0.97. At —20° it is rather hard, at 0° very ductile, at ordinary temperatures like wax, at 50° semi-fluid. It melts at 95°.6 to a mercury-like liquid and yields a colorless vapor at a red heat. It burns with a bright yellow flame when heated in air. Decomposes water energetically, fusing but not taking fire unless the water is warm or the sodium prevented from moving. In other respects acts like potassium.

Used for the preparation of aluminium, magnesium, boron, silicon; in making pure sodium hydroxide and as reducing agent in the laboratory, chiefly in the form of sodium amalgam.

History. Discovered by Davy, 1807, who obtained it by electrolysis of the hydroxide.

COMPOUNDS OF SODIUM.

Sodium Oxides. *Sodium monoxide,* Na_2O, and *peroxide,* Na_2O_2, resemble closely the potassium oxides in their formation and properties.

Sodium Hydroxide, $NaOH$. Prepared in the same way as potassium hydroxide. Like potassium hydroxide in its properties, but somewhat less powerful in its chemical action. Sodium hydroxide (sodium hydrate, caustic soda) is largely used in soapmaking, in the refining of paraffine, petroleum, etc., and in the preparation of water glass. It is also an important reagent in many chemical operations.

Sodium Chloride, NaCl. Occurs as *rock-salt* in immense deposits, and in solution in sea-water (which contains about 2.75 %), salt lakes and saline springs.

Formed by the direct union of its elements and by the action of HCl on NaOH, Na_2CO_3, etc.

Obtained from its natural sources: by quarrying the deposits of rock-salt or by evaporation of its solutions. White cubical crystals. Melts and volatilizes at a red heat. Dissolves in about three times its weight of cold water and is only slightly more soluble in hot water. Ordinary salt is hygroscopic on account of the presence of $MgCl_2$ as impurity.

Sodium chloride or common salt is used as a condiment, in preserving food and wood, in agriculture, in glazing cheap pottery, in several metallurgical operations, in the soap industry, and as the source of chlorine, hydrochloric acid and all sodium compounds.

Sodium Bromide, NaBr, and **Sodium Iodide,** NaI, are usually prepared by adding bromine or iodine to a solution of NaOH, which is then evaporated and heated with carbon (to reduce the bromate or iodate formed) extracted with water and crystallized. **Sodium Fluoride,** NaF, is obtained by neutralizing a solution of HF with Na_2CO_3. All these salts resemble NaCl in their appearance and properties.

Sodium Hypochlorite, NaClO, is known only in aqueous solution, and is used as a bleaching agent under the name of chloride of soda or Labarraque's liquor. It is prepared by leading chlorine into a solution of NaOH. On heating, the hypochlorite decomposes into chloride and chlorate.

Sodium Chlorate, $NaClO_3$, resembles potassium chlorate, but is much more soluble, and hence is obtained pure with greater difficulty. It is used in making aniline black.

Sodium Sulphate, Na_2SO_4, occurs in sea water, in many mineral springs and to some extent as an efflorescence on the soil and in caves.

Prepared on an enormous scale as salt-cake in the first step in Leblanc's soda process (cf. soda), and obtained in the manufacture of nitric acid from sodium nitrate.

Sodium sulphate forms large monoclinic prisms ($Na_2SO_4+10H_2O$) which effloresce in the air, melt in their water of crystallization at 33° and lose all of it below 100°. It dissolves most freely at 34° and readily forms a supersaturated solution, which crystallizes suddenly with considerable rise of temperature when a small crystal of the salt is thrown into it.

Used chiefly in making sodium carbonate; also (instead of caustic soda) in glass making; and in medicine under the name of *Glauber's salt.*

Sodium Sulphite, Na_2SO_3, is obtained by leading sulphur dioxide into a solution of sodium carbonate. If the solution is saturated with sulphur dioxide while hot, *acid sodium sulphite,* $NaHSO_3$, is formed. Both salts are used as an antichlor in bleaching and washing wool. The acid salt is also used for preserving beer, wine, etc.

Sodium Thiosulphate, $Na_2S_2O_3$, (commonly called hyposulphite of soda) is obtained by boiling sodium sulphite with sulphur, and by decomposing the calcium thiosulphate produced by the oxidation of alkali waste from Leblanc's process by sodium carbonate or sulphate. It crystallizes as $Na_2S_2O_3+5H_2O$. The crystals are stable, but the solution gradually decomposes into sulphate and sulphur.

Used largely as antichlor; in photography and the extraction of silver as a solvent of silver chloride; and as a mordant in calico printing.

Sodium Nitrate, $NaNO_3$, occurs in large deposits as *Chili Saltpetre.* Crystallizes in obtuse rhombohedrons, which resemble cubes (hence "cubic saltpetre"). Is hygroscopic and hence unfitted for gunpowder; very soluble in water. A less powerful oxidizing agent than nitre.

Used chiefly for the preparation of potassium nitrate and nitric acid, and as a fertilizer.

Sodium Carbonate, Na_2CO_3, occurs as *natron* and *trona* (sesqui-carbonate) both in the solid form and in solution in mineral waters. Sodium carbonate is also a constituent of the ashes of marine plants.

Obtained from the natural sources in considerable quantities, which, however, now furnish but a small part of the annual production.

Prepared: a. From NaCl by *Leblanc's process,* which consists of the following operations :—

1. *Preparation of the sulphate* (salt cake process). This is usually effected by heating common salt with sulphuric acid :—

$$2NaCl+H_2SO_4=2HCl+Na_2SO_4.$$

Instead of the sulphuric acid, its components are sometimes employed :—

$$2NaCl+SO_2+O(air)+H_2O(steam)=2HCl+Na_2SO_4.$$

The HCl formed as bye-product is absorbed in water.

2. *Conversion of the sulphate into crude soda* (black-ash process). This is accomplished by heating a mixture of the sulphate with chalk and coal. The coal reduces the sulphate to sulphide, which reacts with the chalk to form sodium carbonate and calcium sulphide :—

$$Na_2SO_4+4C=4CO+Na_2S.$$
$$Na_2S+CaCO_3=CaS+Na_2CO_3.$$

3. *Separation of the sodium carbonate* from the insoluble calcium sulphide by lixiviation. The solution which is obtained, always contains some sodium hydroxide from the reaction of $Ca(OH)_2$ (from CaO formed in the black-ash process) and Na_2CO_3, and usually Na_2S and other impurities. It is either evaporated to dryness, yielding an impure *soda-ash ;* or first oxidized and carbonated and allowed to crystallize.

Alkali-maker's waste. For every ton of soda-ash produced, from one and a half to two tons of insoluble residue is formed. This contains all of the sulphur used in the process as sulphuric acid or sulphur dioxide. It gradually decomposes in the air and gives rise to soluble sulphides which contaminate the streams. Hence the recovery of the sulphur is very important for both economic and sanitary considerations. This is now accomplished in many works, a variety of methods being employed.

b. *The ammonia-soda process.* By this process sodium carbonate is made from common salt by the aid of ammonia and carbon dioxide.

1. Into a solution of common salt and ammonia, carbon dioxide is led; this results in the precipitation of acid sodium carbonate:—

$$NaCl+NH_4OH+CO_2=NH_4Cl+NaHCO_3.$$

2. The acid carbonate is readily converted into the normal salt by heat:—

$$2NaHCO_3=Na_2CO_3+H_2O+CO_2.$$

3. The ammonia is recovered by heating the NH_4Cl formed in the first operation with lime

$$2NH_4Cl+Ca(OH)_2=CaCl_2+H_2O+NH_3,$$

and with the CO_2 set free in the second operation is used to repeat the first operation with a fresh quantity of NaCl. The only waste product is $CaCl_2$. Even this loss can be avoided if magnesia is used instead of lime, since magnesium chloride is decomposed by superheated steam into magnesia and hydrochloric acid.

c. *Manufacture of sodium carbonate from cryolite.*

1. Cryolite (Al_2F_6, 6NaF; occurs in large quantities in Greenland), is pulverized, mixed with chalk and heated; soluble sodium aluminate and insoluble calcium fluoride are formed:—

$$Al_2F_6, 6NaF+6CaCO_3=6CaF_2+6CO_2+Al_2(ONa)_6.$$

2. The sodium aluminate is dissolved in water and

carbon dioxide led into the solution. This precipitates aluminium hydroxide and forms sodium carbonate, which is obtained by evaporation of the solution :—

$$Al_2(ONa)_6+3CO_2+3H_2O=Al_2(OH)_6+3Na_2CO_3.$$

Properties. The anhydrous salt or soda-ash is a white mass or powder, which fuses at a red heat. It dissolves in water with evolution of heat, and separates from a concentrated solution at ordinary temperatures in large transparent monoclinic crystals, having the composition $Na_2CO_3+10H_2O$ (soda crystals or washing soda), which effloresce on exposure to air forming a white powder of the mono-hydrate, Na_2CO_3, H_2O. The decahydrate crystals melt at 34°. Several other hydrates are known.

Uses. Sodium carbonate, soda-ash or soda, is used in enormous quantities in soap-making, glass-making and in many other industries, and in the preparation of many sodium compounds.

Hydrogen Sodium Carbonate, *acid sodium carbonate* or *bicarbonate of soda,* $NaHCO_3$, is prepared by passing CO_2 over layers of the decahydrated carbonate. The gas is slowly absorbed and the crystals lose their water of crystallization and transparency :—

$$Na_2CO_3, 10H_2O+CO_2=2NaHCO_3+9H_2O.$$

Much of the impurity of the soda-crystals is removed in the water and the bicarbonate is consequently much purer than the normal salt.

The bicarbonate is much less soluble in water than the normal salt. On heating it is converted into the latter with loss of carbon dioxide and water. In dry air it loses carbon dioxide and is changed into the so-called sesquicarbonate, Na_2CO_3, $2NaHCO_3$, $2H_2O$.

The bicarbonate is used in medicine, in preparation of baking powders and of effervescing powders and drinks.

Sodium Pyroborate or **Borax,** $Na_2B_4O_7$, *occurs* in certain lakes and in boronatrocalcite ($Na_2B_4O_7 + 2CaB_4O_7 + 18H_2O$). *Obtained:* (a) from the water of borax lakes by evaporation, forming "tincal;" (b) from boronatrocalcite by reaction with sodium carbonate or by precipitation of calcium sulphate by sulphuric acid and neutralization of the boric acid thus formed with sodium carbonate; (c) in large quantities from native boric acid by addition of sodium carbonate.

Properties. Borax crystallizes from solution under ordinary conditions with ten molecules of water of crystallization in large transparent monoclinic prisms. When heated, these swell up, lose water and form a spongy mass of "burnt borax," which melts at a red heat to transparent "borax glass." Fused borax glass dissolves many metallic oxides which impart characteristic colors to the glass.

Used for cleaning metallic surfaces for soldering, in blowpipe analysis and as a flux in metallurgy; in making certain kinds of glass (Strass); glazing earthen ware; for fixing mineral mordants; in washing and for a number of minor purposes.

Sodium Silicate, $Na_2Si_4O_9$. Commonly known as soluble glass. Prepared by heating sand, soda-ash and charcoal, or dissolving powdered flint in a hot concentrated solution of caustic soda.

A glass-like mass which dissolves slowly in boiling water. Used always in solution and for the greatest variety of purposes :—for making artificial stone and hardening building stone; mixed with chalk or zinc oxide as cement; for fixing fresco colors in stereochromy; as mordant in dyeing; as an addition to soap, etc.

Sodium Phosphates. Of the three orthphosphates—Na_3PO_4, Na_2HPO_4 and NaH_2PO_4—the second, *disodium phosphate*, is the most important. It is prepared by adding sodium carbonate to phosphoric acid. Crystallizes in large transparent prisms—$Na_2HPO_4+12H_2O$—which effloresce in the air, lose all their water at 100°; melt at 300° and are converted at a higher temperature into *sodium pyrophosphate*, $Na_4P_2O_7$. Disodium phosphate is used in the laboratory and in medicine.

Sodium Metaphosphate, $NaPO_3$, is formed by heating monosodium orthophosphate, NaH_2PO_4, or microcosmic salt, $HNaNH_4PO_4$.

Sodium Sulphides closely resemble the corresponding potassium compounds.

LITHIUM. $Li^I=7$.

Occurs in small quantities in petalite, lepidolite and spodumene (as silicate); in triphyline (as phosphate); in certain mineral waters; is almost universally diffused in minute quantities in most natural waters and soils, and is found in plants and animals.

Prepared from its chloride by electrolysis.

Properties. Resembles sodium; but is lighter (lightest solid known, sp. gr. 0.59), harder, less fusible (melts at 180°) and oxidizes less readily. Ignites at 200° and burns with a bright white light; decomposes water, but without fusion or ignition.

COMPOUNDS OF LITHIUM.

The salts of lithium resemble those of sodium. The carbonate and phosphate are much less soluble than the corresponding sodium compounds. The bromide, carbonate, citrate and salicylate are employed in medicine. All lithium salts impart an intense crimson color to a non-luminous flame and can be detected in minute traces by the spectroscope.

RUBIDIUM AND CAESIUM.

Rubidium, $Rb^{I}=85.4$. *Occurs* widely distributed in many minerals, waters and plants, but in minute quantities, generally accompanying the other alkali metals.

Prepared in the same way as potassium.

Properties. Closely resembles potassium, but is heavier (sp. gr. 1.52), melts at a lower temperature (38°.5), yields at a red heat a blue vapor, and is more active chemically. Its salts are like those of potassium. They give a red flame coloration.

Caesium, $Ce^{I}=133$. *Occurs* in the rare mineral pollux and in some mineral waters, usually with rubidium and potassium, but is not taken up by plants. It is still more rare than rubidium.

Prepared by electrolysis of a melted mixture of the cyanides of caesium and barium.

Properties. In general like rubidium, but melts lower (26–27°), is heavier (sp. gr. 1.88). It ignites quickly in the air and is the most electro-positive element known. Its salts resemble those of potassium and give a bright blue flame coloration.

Both metals were discovered by Bunsen by spectrum analysis (Caesium 1860, Rubidium 1861), caesium being the first element discovered by this means.

COMPOUNDS OF AMMONIUM.

The monad group $NH_4{}^I$, although incapable of isolation, occurs in a number of important salts which bear strong analogies to those of the alkali metals. It is also probably present in *ammonium amalgam* (a spongy mass formed when sodium amalgam is placed in a solution of ammonium chloride), which rapidly decomposes into mercury, ammonia and hydrogen.

The ammonium salts are all volatile and decomposed by heat. They yield ammonia by the action of bases.

Ammonium Chloride or *Sal-ammoniac,* NH_4Cl, *occurs* as white crust or efflorescence in many volcanic regions and is occasionally found about ignited coal seams.

Prepared: (a) from the ammoniacal liquor of gas works by leading the ammonia produced by heating this with lime into hydrochloric acid. (b) From ammonium sulphate by the action of sodium chloride :—

$$(NH_4)_2SO_4+2NaCl=Na_2SO_4+NH_4Cl.$$

The Glauber's salt is easily separated on account of its less solubility. In both cases the solution of ammonium chloride is evaporated, and the salt purified by sublimation, forming a tough, fibrous, semitransparent mass; or when the vapor is quickly cooled a light crystalline powder.

Properties. Colorless salt of sharp saline taste. Sublimes without melting and separates completely at 350° into NH_3 and HCl, which unite again on cooling. Very soluble in water with absorption of much heat. The solution on boiling loses a little NH_3 and becomes acid.

Used in medicine, for the preparation of ammonia and ammonium compounds, in soldering and tinning to give clean metallic surfaces, in dyeing, in galvanic batteries and as a valuable reagent in the laboratory.

Ammonium Sulphate, $(NH_4)_2SO_4$, occurs in small quantities in certain volcanic regions. Prepared by leading the ammonia from ammoniacal gas liquor into sulphuric acid.

Very soluble in water ; melts at 140° and decomposes at a higher temperature. Used in large quantities as fertilizer, in the preparation of ammonia alum and other ammonia compounds.

Ammonium Nitrate, NH_4NO_3, prepared from ammonium hydroxide or carbonate by neutralization with nitric acid. Crystallizes in rhombic prisms. Deliquescent and exceedingly soluble in water with great absorption of heat. Decomposes on heating into water and nitrogen monoxide and is used for the preparation of this gas, and in freezing mixtures.

Ammonium Nitrite, NH_4NO_2, best prepared pure by double decomposition of silver nitrite and ammonium chloride, or of lead nitrite and ammonium sulphate An indistinctly crystallized mass which decomposes on heating into water and nitrogen.

Ammonium Carbonates. *Commercial carbonate* or *sesquicarbonate of ammonium*, sal-volatile, or salt of hartshorn, is obtained by sublimation from a mixture of ammonium chloride or sulphate and chalk. It is a compound of acid ammonium carbonate and ammonium carbamate : $(NH_4)HCO_3+NH_4CO_2NH_2$. It has a strong ammoniacal smell and pungent caustic taste. Decomposes slowly in the air leaving a white powder of $(NH_4)HCO_3$. Used in smelling salts, baking powders, and in medicine. *The normal carbonate* $(NH_4)_2CO_3$ is unstable, losing ammonia and passing into the acid carbonate, NH_4HCO_3, which at 38° decomposes into CO_2, NH_3 and H_2O.

Ammonium Phosphates are like those of sodium and potassium. The most important is the *hydrogen sodium ammonium phosphate* or microcosmic salt, $HNa(NH_4)PO_4+4H_2O$, which is converted on heating into a transparent glass of sodium metaphosphate and when fused dissolves many metallic oxides with characteristic colors. The salt is used on this account in blowpipe analysis.

Ammonium Sulphide, $(NH_4)_2S$, and *Hydrosulphide.* NH_4SH, are formed in the same way as the corresponding potassium compounds, which they resemble in their properties. These solutions decompose slowly on standing, becoming yellow from the formation of polysulphides, and finally colorless with precipitation of sulphur.

METALS OF THE ALKALINE EARTHS.

CALCIUM. STRONTIUM. BARIUM.

The oxides of these metals, known as alkaline earths, were supposed to be elementary substances until Davy, 1808, showed that each is a compound of oxygen and a metal. They are all prepared by electrolysis of their fused chlorides (calcium sometimes by reduction of its chloride or iodide with sodium).

Properties. The metals are slightly harder than lead and malleable; they fuse at or above a red heat; cannot be volatilized; oxidize readily in moist air and decompose cold water energetically. They are not used for any purpose. The *hydroxides* are all more or less soluble in water, giving alkaline solutions. The *sulphates* are insoluble or very slightly soluble; the normal carbonates are insoluble.

CALCIUM. $Ca^{II}=40$.

Occurs only in combination, chiefly as carbonate, sulphate, phosphate, silicate and fluorides. Calcium is one of the most abundant elements and its salts are widely distributed forming a large portion of the crust of the the earth; and existing in most natural waters and in all plants and animals.

Properties. A yellow metal, sp. gr. 1.6. Burns with a yellow flame. Is less strongly electro-positive than strontium and barium.

COMPOUNDS OF CALCIUM.

Calcium Oxide, or quicklime, CaO, is formed from calcium carbonate by heating:—$CaCO_3=CO_2+CaO$, and is prepared on a large scale by "burning" limestone in kilns. A white amorphous infusible mass. In air it absorbs moisture, becoming "air slaked" and crumbling into a sandy powder of hydroxide which is gradually converted into carbonate by carbon dioxide. Unites with water with great heat, forming calcium hydroxide:—

$$CaO+H_2O=Ca(OH)_2.$$

Used for drying gases and liquids, and for producing the lime light; chiefly employed in the form of hydroxide or slaked lime.

Calcium Hydroxide, *slaked lime,* $Ca(OH)_2$, prepared from quicklime by adding about one-third its weight of water (slaking).

A white soft powder. Slightly soluble in water (lime water), but less soluble in hot water than in cold. The solution has a feebly caustic taste and alkaline reaction; both solution and powder absorb carbon dioxide with the formation of calcium carbonate. Decomposes at a red heat into CaO and H_2O. *Milk of lime* is calcium hydroxide suspended in water.

Calcium hydroxide is used in making mortar; in the preparation of the caustic alkalies, ammonia, "chloride of lime," and for a great variety of technical purposes.

Mortars. *Ordinary mortar* is a stiff paste of slaked lime, sand and water. The sand is necessary to prevent shrinking when the mortar hardens. When exposed to the air, mortar first loses water and "sets," and then "hardens" from the gradual formation of calcium carbonate by absorption of carbon dioxide from the air.

Hydraulic Mortar. Lime prepared from limestone which contains more than 10 per cent. of silica, has the property of hardening under water. Such a lime may also be obtained by heating a mixture of ordinary lime and clay. *Portland cement* is an hydraulic mortar prepared by burning a mixture of finely ground chalk and clay.

Calcium Chloride, $CaCl_2$, occurs in solution in sea-water and many mineral springs and is a constituent of a few minerals. Prepared by dissolving calcium carbonate (chalk) in hydrochloric acid. It is obtained in large quantities as bye-product in several chemical industries:—e. g. ammonia soda process.

Properties. Crystallizes from solution in (a) large hexagonal prisms as $CaCl_2+6H_2O$, which dissolve in water with absorption of much heat; melt, on heating, in their water of crystallization; lose four molecules of it at 200°, forming (b) a white porous mass which melts at red heat, becomes anhydrous and solidifies on cooling to (c) a white crystalline mass, soluble in water with evolution of heat. In all three forms the salt is very deliquescent and is used for drying gases and liquids. (It cannot be used for drying NH_3 as this is absorbed by it with the formation of a voluminous powder, $CaCl_2+8NH_3$). It is also employed in making freezing mixtures and for some other minor purposes.

Chloride of Lime, Bleaching Powder, apparently $CaOCl_2$, is made on an immense scale in alkali works by passing chlorine over thin layers of dry slaked lime. It is a white powder of faint chlorine like odor. Becomes moist in the air and gradually decomposes with absorption of carbon dioxide. Dissolves more or less completely in water. Evolves chlorine readily with the weakest acids and is used extensively as bleaching and disinfecting agent.

Calcium Fluoride, CaF_2, occurs as fluor spar, in apatite and in minute quantities in bones, enamel of the teeth, ashes of certain plants and in the sea and certain mineral waters. May be prepared by precipitation of a solution of a calcium salt by a soluble fluoride. Melts easily at a red heat and has been used from early times as a flux in metallurgy. It serves also for the preparation of hydrofluoric acid, and fine specimens of fluor spar have been made into vases and other ornamental objects.

Calcium Sulphate, $CaSO_4$, occurs as anhydrite; more abundantly as gypsum, $CaSO_4+2H_2O$; and is found in the soil and natural waters (causing "permanent hardness"). Is precipitated from calcium salts by sulphuric acid or soluble sulphates. Very slightly soluble in water. Loses most of its water of crystallization at 120° and is thus converted into "burnt gypsum" or "plaster of paris," which, when mixed with water to the consistance of cream, combines with it and hardens. Heated above 200°, gypsum become "dead burnt" and hardens very slowly.

The variety of gypsum called alabaster is used for vases and various ornaments; the coarser kinds are employed as fertilizer. Plaster of paris is used as a cement, for ornamental plaster work and for making plaster casts. The artificial salt is used in paper manufacture as a filling for paper.

Calcium Phosphates. *Normal calcium phosphate*, $Ca_3(PO_4)_2$, occurs in apatite and phosphorite and is the chief inorganic constituent of bones. It is insoluble in water, but dissolves readily in dilute acids and in solutions of certain salts. It is not decomposed by heat.

Bone-ash is used for making phosphorus, and bone-ash and natural phosphates for the manufacture of "superphosphate of lime" (a mixture of calcium sulphate and acid phosphate, $H_4Ca(PO_4)_2$, obtained by treating normal phosphate with dilute sulphuric acid), which is extensively used as a fertilizer.

Calcium Carbonate, $CaCO_3$, occurs very abundantly as limestone, chalk, marble, and with magnesium carbonate in dolomite; forms the greater part of egg-shell, coral, etc.; and is found in solution in nearly all waters, causing "temporary hardness."

Insoluble in pure water, but dissolves slightly in water containing free CO_2, probably as acid carbonate $CaH_2(CO_3)_2$. At a red heat it loses CO_2 and is converted into CaO.

Transparent crystals of calcite are used in optical instruments; marble, as building stone, for statuary, etc. ; chalk, in the manufacture of glass, soda, etc. ; limestone, chiefly for making quicklime, CO_2 and as flux.

Calcium Silicates occur in almost all natural silicates, and form an important constituent of glass.

Glass is an amorphous mixture of various silicates, especially those of the alkalies and alkaline earths. The chief materials are:—(1) Silica : in the form of white sand, ignited flint or quartz. (2) Alkali : potash, soda or sodium sulphate. (3) Calcium carbonate in one of its natural forms. These are fused together.

a. *Bohemian Glass* is a silicate of potassium and calcium. It is very difficultly fusible ; of all glass is least acted on by chemical reagents. Used for chemical apparatus.

b. *Window or Crown-Glass* is a silicate of sodium and calcium. More fusible and more readily at-attacked, but harder than the bohemian glass.

c. *Bottle Glass,* or *Common Green Glass,* is a silicate of sodium and calcium containing alumina and oxide of iron.

d. *Flint Glass,* or *Crystal* is a potash-lead silicate.

Glass fuses at high temperatures, passing through all gradations of viscosity on cooling. It is nearly insoluble in water and acids and differs in this respect from its constituent silicates,as well as in its non-crystalline structure ; for the alkaline silicates, though non-crystalline, are soluble in water, and calcium silicate, though insoluble in water, forms a crystalline mass which is soluble in acids.

DETECTION AND ESTIMATION OF CALCIUM.

Calcium compounds give an orange red flame coloration. Calcium is usually precipitated as oxalate, which is converted into carbonate or oxide by heat and weighed in one of these forms. It is occasionally determined as sulphate.

STRONTIUM. $Sr^{II}=87.2$.

Occurs as sulphate (celestine) and carbonate (strontianite). It is a yellow metal. Sp. gr. 2.5.

Its Compounds closely resemble those of calcium. They are obtained both from celestine and strontianite.

a. Celestine is heated with carbon and thus converted into sulphide :— $SrSO_4+2C = 2CO_2+SrS$; which is then changed into hydroxide by the action of superheated steam :—$SrS+2H_2O=Sr(OH)_2+H_2S$; or into chloride, nitrate, etc. by the corresponding acid :—e. g. $SrS+2HCl=SrCl_2+H_2S$.

b. Strontianite gives the oxide on heating, and the salts may be obtained by the action of acids on this or on the mineral itself.

The most important strontium compounds are :— The *hydroxide* $Sr(OH)_2$, which is used in sugar-making to extract the sugar from molasses; and the *nitrate,* which is used as a constituent of "red fire."

Strontium salts color the non-luminous flame a magnificent crimson.

BARIUM. $Ba^{II}=137$.

Occurs chiefly as sulphate (heavy-spar) and carbonate (witherite). The metal resembles strontium and calcium. It is the most strongly positive member of the group.

COMPOUNDS OF BARIUM.

The salts of barium are obtained by dissolving witherite in acids, or from heavy spar after reduction to sulphide (cf. Strontium). Barium salts resemble those of strontium and calcium, but are distinguished from them by being very poisonous.

Barium Oxide, BaO, is prepared best by heating barium nitrate. It can also be made from the carbonate by heating it with carbon. It slakes in water like lime forming

Barium Hydroxide, $Ba(OH)_2$, which may also be prepared by the action of steam on the carbonate at a red heat, or by precipitation of a concentrated solution of a barium salt with an alkaline hydroxide.

It is much more soluble than calcium hydroxide and crystallizes as $Ba(OH)_2+8H_2O$. Unlike calcium and strontium hydroxides, barium hydroxide is not converted by heat into its oxide. Its solution (caustic baryta) is more strongly alkaline than those of the other alkaline earths. Used as reagent in the laboratory.

Barium Dioxide, BaO_2, is formed when BaO is heated to redness in a current of dry air or oxygen. Decomposes at a bright red heat into BaO and oxygen. Dissolves in dilute acids with the formation of hydrogen peroxide and is used for its preparation. With concentrated sulphuric acid it yields ozone.

Barium Chloride, $BaCl_2$, crystallizes as $BaCl_2+2H_2O$. Quite soluble in water ; insoluble in alcohol and concentrated hydrochloric acid. Used as reagent in the laboratory; for preventing boiler incrustations; as an antiseptic and for the preparation of "permanent white."

Barium Sulphate, $BaSO_4$, occurs as heavy spar. Prepared by precipitation of $BaCl_2$ with sulphuric acid. Both the natural sulphate and that formed by precipitation are insolnble in water and acids. Heated with carbon, barium sulphate is reduced to sulphide.

Finely ground heavy-spar both alone and mixed with more or less white lead, is largely used as a paint; the artificial sulphate is also made on a large scale for this purpose, being known as *permanent white* or *blanc fixe.* Heavy spar is the chief source of the artificial barium compounds.

Barium Nitrate, $Ba2NO_3$, is largely used in pyrotechny, especially for the preparation of "green fire."

Barium Carbonate, $BaCO_3$, occurs as witherite. Artificially prepared by precipitation, it is a white amorphous insoluble powder. Melts at a white heat and loses CO_2 slowly; heated with carbon it forms BaO. Witherite is used for making other barium compounds and as rat bane. The artificial carbonate is used in chemical analysis.

DETECTION AND ESTIMATION OF BARIUM.

Barium salts give a yellowish green flame coloration, and the insoluble sulphate is precipitated from their solutions by sulphuric acid. Barium is almost always weighed as sulphate, occasionally as carbonate.

METALS OF THE MAGNESIUM GROUP.

MAGNESIUM. BERYLLIUM. ZINC. CADMIUM.

These four metals are all dyads. They can be volatilized and burn when heated in air with a bright flame. They do not decompose water at ordinary temperatures. Their oxides, hydroxides and carbonates are insoluble in water (except magnesium hydroxide which is slightly soluble). The hydroxides and carbonates are converted into oxides by heat.

MAGNESIUM. $Mg^{II}=24$.

Occurs abundantly, but only in combination, chiefly as carbonate, silicate, chloride and sulphate. Its soluble salts are found in most natural waters.

Prepared from its chloride by electrolysis or reduction with sodium.

Properties. Silver-white metal of low specific gravity (1.75); melts at a red heat and boils about 1020°. Moderately hard; malleable and ductile, and is usually rolled into ribbon or drawn into wire.

In moist air it becomes covered with a white film of hydroxide; decomposes water somewhat at 100°. Ignited, it burns with an intense white light, which is very rich in actinic rays. Readily soluble in dilute acids; insoluble in caustic alkalies. Is a strong reducing agent.

Used to furnish an artificial light for photography and signalling (magnesium light), and as a reducing agent in the laboratory.

COMPOUNDS OF MAGNESIUM.

Magnesium Oxide or *Magnesia,* MgO, occurs as periclase. Prepared by heating the carbonate.

Very light white powder of feeble alkaline taste. Infusible except at the heat of the oxyhydrogen blowpipe. Insoluble in water, but combines with it forming $Mg(OH)_2$.

Magnesium oxide, *magnesia usta,* or "calcined magnesia," is used in medicine as an antacid, etc. On account of its infusibility it serves as material for crucibles. Heavy or dense magnesia—obtained by heating the chloride or nitrate—hardens under water as $Mg(OH)_2$. A mixture of magnesia and chalk, and burnt dolomite (a mixture of magnesia and lime) have the same property, and are used as hydraulic cements, and for making artificial stone.

Magnesium Hydroxide, $Mg(OH)_2$, occurs as brucite. Obtained as a white precipitate on adding caustic potash or soda to a magnesium salt and is formed when magnesia is moistened with water. Soluble in 5500 parts of water.

Magnesium Chloride, $MgCl_2$, occurs in sea-water, and salt springs and salt beds, notably in double salts in the Stassfurt deposits, whence it is obtained in large quantities. It is very soluble in water and is obtained by evaporation in deliquescent crystals, $MgCl_2+6H_2O$. On heating, this hydrated salt evolves H_2O and HCl and is converted into oxide.

The *anhydrous salt* is prepared by heating ammonium magnesium chloride, $MgCl_2,NH_4Cl+6H_2O$. It is deliquescent, dissolves in water with evolution of heat, melts on heating and can be distilled.

Magnesium chloride forms other double salts like that with ammonium, the most important being carnallite, $MgCl_2,KCl+6H_2O$.

Used for dressing cotton and woollen goods and for some other purposes. Strongly ignited magnesia mixed with a concentrated solution of the chloride, hardens after a few hours (Sorel's cement).

Magnesium Sulphate, $MgSO_4$, occurs as kieserite, $MgSO_4+H_2O$, and as Epsom Salt, $MgSO_4+7H_2O$. Kieserite is very difficultly soluble in water; with a little water it forms a hard coherent mass, which slowly dissolves in more water as $MgSO_4+7H_2O$. Epsom salt is found in many mineral waters and as an efflorescence in caves, etc.

Epsom salt has an unpleasant, bitter taste; is readily soluble in water, from which it crystallizes at ordinary temperatures as $MgSO_4+7H_2O$. Forms double salts with ammonium and potassium, e. g. $MgSO_4,K_2SO_4+6H_2O$.

Used in medicine, as a dressing for cotton goods; in aniline dyeing; as cheap sulphate in the preparation of Glauber's salt, permanent white and other sulphates; as reagent in the laboratory and as fertilizer.

Magnesium Carbonate. *Normal magnesium carbonate*, $MgCO_3$, occurs as magnesite and with calcium carbonate in dolomite ($CaCO_3,MgCO_3$). Slightly soluble in water containing CO_2 as acid carbonate (cf. calcium carbonate) and hence is found in many natural waters. On boiling this solution or on adding potassium or sodium carbonate to a solution of a magnesium salt, a *basic magnesium carbonate* is precipitated, which forms on drying a voluminous powder, "magnesia alba," $Mg_4(OH)_23CO_3$.

Magnesia alba is insoluble in water, but dissolves in a solution of ammonium salts. It is used in medicine and in the preparation of tooth powders and cosmetics.

Ammonium Magnesium Ortho-Phosphate, $MgNH_4PO_4+6H_2O$, is precipitated from a magnesium solution on adding ammonium salts and a phosphate. It is a crystalline powder insoluble in ammonia. On heating, it loses water and ammonia and is converted into magnesium pyrophosphate, $Mg_2P_2O_7$. It serves for the determination of magnesium and of phosphoric acid.

Magnesium Silicates form a number of minerals, of which the more important are talc (soapstone) serpentine and meerschaum. Hornblende (asbestos) is a double silicate of magnesium and calcium.

DETECTION AND DETERMINATION OF MAGNESIUM.

Magnesium is detected by the formation of ammonium magnesium phosphate and weighed as magnesium pyrophosphate.

BERYLLIUM. $Be^{II}=9$.

Occurs in a number of minerals, especially in beryl (a double silicate of aluminium and beryllium). The metal is prepared by reduction of its chloride with sodium. In a compact state, it oxidizes when heated only on the surface ; in a finely divided state it burns with great brilliancy to BeO. Melts a little lower than silver ; dissolves in dilute hydrochloric or sulphuric acid and in caustic potash or soda.

COMPOUNDS OF BERYLLIUM.

Beryllium salts resemble those of magnesium. The soluble salts have a sweetish taste and on this account the oxide was at first called *glucina,* and the metal *glucinum.*

ZINC. $Zn^{II}=65$.

Occurs in combination as oxide (zincite and in franklinite), sulphide (blende), carbonate (smithsonite), and silicate (calamine). All of these minerals serve as ores of zinc.

Extraction. The ores are oxidized by roasting and the oxide reduced by coal. The zinc distils, is condensed in a receiver and purified by redistillation.

Properties. Bluish-white metal, harder than silver. Specific gravity, 6.9. Brittle at ordinary temperatures, it becomes highly malleable and ductile between 100° and 150° ; at 200° is so brittle that it can be powdered in a mortar. If rolled into sheets while malleable, it retains its malleability when cooled. It has the highest coefficient of expansion of any useful metal. Melts at 433° and boils at 1040. In moist air it becomes coated with a film of basic carbonate. Heated in air it burns with an intense bluish-white light to ZnO. Decomposes water only at a red heat. Readily soluble in dilute acids and dissolves on boiling in caustic alkalies. Reduces many metals from their solutions.

Zinc dust, obtained with the compact metal in the process of distillation, is a fine grey powder consisting of metallic zinc and some oxide. It is a powerful reducing agent, decomposing water at ordinary temperatures.

Uses. Zinc is used in electrical batteries, for castings and in the form of sheets for roofing and other purposes. It is also used as a constituent of brass and other alloys, for coating ("galvanizing") iron, desilverizing lead, as a source of zinc preparations and in the laboratory for making hydrogen.

COMPOUNDS OF ZINC.

Zinc Oxide, occurs as zincite and (with iron and manganese) in franklinite. It is prepared on a large scale by burning zinc vapor; for pharmaceutical purposes by heating basic carbonate. A white soft powder; becomes yellow when heated and white again on cooling. Insoluble in water; soluble in acids and caustic alkalies.

Used as paint (zinc-white) instead of white lead, and for some minor purposes; also in medicine.

Zinc Hydroxide, $Zn(OH)_2$, is precipitated from zinc salts by caustic alkalies, dissolving readily in an excess of the reagent forming a zincate:

$$Zn(OH)_2+2NaOH=2H_2O+(NaO)_2Zn.$$

Zinc Chloride, $ZnCl_2$, is obtained by heating zinc in a current of chlorine, or zinc sulphate with sodium chloride. In solution it is prepared on a large scale by dissolving zinc or blende (ZnS) in hydrochloric acid. Anhydrous zinc chloride is a white soft mass (butter of zinc). It is very deliquescent, dissolves freely in water and withdraws it or its elements from many organic compounds. On evaporation of its solution it is partly convertedinto oxide (cf. $MgCl_2$). Melts a little above 100° and distils at a red heat.

Used largely for preserving timber, to withdraw water from various substances, for separating wool from cotton, in the manufacture of aniline colors, as mordant in cotton printing, as soldering fluid, as disinfectant and as a caustic in surgery.

Zinc oxide mixed with a concentrated solution of zinc chloride, gives a plastic mass, which soon hardens from the formation of an oxychloride, and is used as a cement.

Zinc Sulphate, or *white vitriol,* $ZnSO_4$, occurs in zinc mines, formed by the oxidation of blende, and is made on a large scale by oxidizing blende partially by roasting. It crystallizes as $ZnSO_4+7H_2O$. On heating, it melts in its water of crystallization, loses 6 mol. at 100°, the last at a higher temperature. When heated more strongly it is decomposed into oxide, sulphur dioxide and oxygen. Used as "dryer" in oil paints and varnishes, as mordant, preservative and disinfectant, and in medicine.

Zinc Sulphide, ZnS, occurs as blende, which is one of the most important ores of zinc and is also used in making sulphuric acid and zinc sulphate. Zinc sulphide is obtained as a white precipitate on adding ammonium sulphide to a zinc solution. It is soluble in dilute acids, except acetic, and hence is precipitated from an acetic acid solution by hydrogen sulphide.

DETECTION AND DETERMINATION OF ZINC.

The white sulphide is the most characteristic compound of zinc and may serve for its detection. Zinc is usually weighed as oxide.

CADMIUM. $Cd^{II}=112$.

Occurs as sulphide in the rare mineral greenockite, and in certain ores of zinc. In the process of zinc smelting, cadmium vapor comes off with the first portions of zinc and collects in the zinc dust as oxide. It is *prepared* from this by distillation with coal.

Properties. Tin-white metal of fibrous fracture. Soft, malleable and ductile. Melts at 315°, boils at 860°, yielding a dark yellow vapor. Tarnishes in moist air. Burns when heated with a brown smoke of CdO. Soluble in dilute acids.

Used in making fusible alloys and cadmium preparations, and in an amalgam for stopping teeth.

COMPOUNDS OF CADMIUM.

Cadmium salts resemble those of zinc. The soluble salts are poisonous. Zinc precipitates cadmium from its solutions.

Cadmium Sulphide, CdS, occurs as greenockite. Prepared by precipitating cadmium solutions with H_2S or alkali sulphide. It is a yellow powder used as pigment (cadmium yellow or jaune brilliant) and in pyrotechny for producing a blue fire.

METALS OF THE LEAD GROUP.

LEAD. **THALLIUM.**

Both metals are soft, of high specific gravity and their basic oxides have an alkaline reaction. Lead is in many respects analogous to the metals of the alkaline earths, while thallium in many of its properties resembles the alkali metals; they differ however from both of these groups in the formation of black sulphides insoluble in water, and sparingly soluble chlorides.

LEAD. Pb (Plumbum)=206.4.

Occurs rarely native; chiefly combined as sulphide (galena); also as carbonate, sulphate, phosphate, chromate and other salts.

Obtained almost entirely from galena by processes which differ according to the purity of the ore :—

1. *Air reduction process* for comparatively pure ores. The galena is roasted until part of the sulphide is converted into oxide and sulphate; the temperature is then raised and the following reactions take place :—

$$PbS+2PbO=SO_2+3Pb.$$
$$PbS+PbSO_4=2SO_2+2Pb.$$

2. *Carbon reduction process* for less pure ores, which consists in roasting the ore to oxide, and reducing by coal.

3. *Precipitation process* for ores containing sulphides of other metals. The galena is reduced by heating with iron :—

$$PbS+Fe=FeS+Pb.$$

Desilverization.—The lead obtained by any of these processes always contains silver. When the silver amounts to eight ounces or more per ton, it can be profitably extracted by removing the lead by oxidation (cupellation). If the amount of silver is less than this, it may be concentrated (a) by allowing the melted metal to cool slowly and removing the crystals which form by means of a perforated ladle. These crystals consist of almost pure lead, while the still liquid lead contains most of the silver ; or (b) by adding a small quantity of zinc to the fused lead and allowing it to crystallize as before. In this case a silver-zinc alloy is formed containing nearly all the silver and crystallizing first. The small amount of zinc which remains in the lead is removed by oxidation with superheated steam.

Properties. Bluish metal, soft and malleable but of small tenacity. Specific gravity 11.37. Melts at 335° and is somewhat volatile at a high heat. Tarnishes in the air from the formation of a film of suboxide. Heated in air, it is gradually converted into lead-ash, a yellowish-gray mixture of suboxide and monoxide, which on longer heating is wholly changed into monoxide. In contact with pure water and air, lead is slowly oxidized to $Pb(OH)_2$ which is somewhat soluble. In water containing carbonates and sulphates it is scarcely attacked ; while water containing certain other salts, notably ammonium nitrate, dissolves lead to some extent. Soluble in nitric acid; hardly acted on by sulphuric or hydrochloric acid.

Used as sheet-lead for roofs, sulphuric acid chambers, etc.; as lead-foil for packing; for water pipes, shot and bullets, making many alloys and lead preparations, and also for the extraction of silver and gold.

COMPOUNDS OF LEAD.

Lead is a dyad in all of its compounds except PbO_2 and certain organic compounds in which it acts as a tetrad. The salts are mostly colorless and only a few of them are soluble; these have a sweetish astringent taste and together with the oxides and carbonates are poisonous. Basic salts are readily formed, most of which are insoluble.

Lead Oxides. *Lead Suboxide,* Pb_2O, is formed by heating lead oxalate to 300°: $2PbC_2O_4 = CO + 3CO_2 + Pb_2O$. It is a dark grey powder which, when heated in absence of air, decomposes into Pb and PbO.

Lead Monoxide, PbO, obtained by heating lead, its carbonate or nitrate. It is a yellow powder called "massicot." Fuses at a red heat yielding "litharge," which is also obtained as bye-product in the extraction of silver by cupellation. Lead oxide is strongly basic and absorbs CO_2 from the air. Dissolves in 7000 pts. of water as $Pb(OH)_2$, giving a solution of alkaline reaction. Like other strong bases it forms a soap with fats, the so-called lead-plaster. Melted with silica or silicates it forms an easily fusible glass.

Litharge is largely used in making certain varieties of glass and for glazing earthen ware ; as a flux in glass and porcelain staining; in the preparation of varnishes, putty, lead-plaster, red-lead and acetate of lead.

Red-Lead or *Minium*, Pb_3O_4, is prepared by carefully heating finely divided massicot or white lead. It is a scarlet powder which is used as a paint, in making lead glass, cement for gas and steam pipes, etc.

Lead peroxide, PbO_2, is obtained by treating red-lead with dilute nitric acid. It is a brown powder, which decomposes on heating into oxygen and monoxide; It is a strong oxidizing agent and is used as such in the mixture for match heads. Lead peroxide behaves as an acid anhydride toward strong bases, uniting with them to form *plumbates*, derived from the corresponding hydroxide $PbO(OH)_2$ or *plumbic acid.*

Lead Chloride, $PbCl_2$, occurs as cotunnite. Formed from lead by the action of boiling hydrochloric acid or warm aqua regia ; by dissolving galena, lead oxide or carbonate in hydrochloric acid ; and is precipitated on adding hydrochloric acid to a lead solution. Soluble in 135 parts of cold water ; in 30 parts at 100° and crystallizes from the hot solution on cooling heavy white silky needles. Easily fusible ; non-volatile. Unites with lead oxide or hydroxide to form various oxychlorides, which are white or yellowish according to the proportion of oxygen they contain, and are used as pigments (Turner's yellow, Cassel yellow, Pattinson's white-lead).

Lead Sulphate, $PbSO_4$, occurs as anglesite or lead vitriol. Is formed as a heavy white precipitate on adding sulphuric acid to a lead solution. Almost insoluble in water; quite insoluble in dilute sulphuric acid and in alcohol. Dissolves slightly in hot concentrated sulphuric acid and is precipitated when this is diluted. Readily soluble in ammoniacal salts, especially in the acetate and tartrate.

Lead Nitrate, $Pb2NO_3$, forms hard octohedrons readily soluble in water; sparingly soluble in dilute and insoluble in concentrated nitric acid. Melts at a red heat and decomposes into lead oxide, oxygen and NO_2. The aqueous solution dissolves lead oxide on boiling, forming a sparingly soluble basic nitrate, $HOPbNO_3$.

Lead nitrate is largely used in dyeing and calico-printing and for the preparation of chrome-yellow.

Lead Carbonates. *Normal Lead Carbonate,* occurs as cerusite and is precipitated from a lead solution by ammonium carbonate. Sodium and potassium carbonates precipitate various basic carbonates according to the concentration and temperature of the lead solution.

White Lead consists of basic lead carbonates—chiefly $(HO)_2Pb_32CO_3$—and is prepared on a large scale (a) by simultaneous action of acetic acid vapors and carbon dioxide upon lead (Dutch and German methods); (b) by leading carbon dioxide into a solution of basic lead acetate (French and English methods); (c) by grinding litharge (or any insoluble basic lead salt) with water and bicarbonate of soda.

White lead is a heavy white powder, which is very extensively used for paint in preference to zinc-white or permanent-white and in spite of its poisonous character and blackening by sulphurretted hydrogen, because of its great "covering power."

Lead Sulphide, PbS, occurs as galena the chief ore of lead, and is precipitated as a black powder from lead solutions by hydrogen sulphide. Converted by nitric acid partly into nitrate and partly into sulphate; the proportion of the latter being greater the stronger the acid employed.

DETECTION AND DETERMINATION OF LEAD.

Lead is readily detected by means of its chloride, sulphate, chromate or sulphide.

It is usually weighed as sulphate.

THALLIUM. $Th^{I \text{ and } III} = 203.7$.

Occurs in the rare mineral crookesite and in small quantities in many pyrites.

It is best *prepared* from the flue dust of sulphuric acid works in which pyrites containg it is burned, by boiling the dust with dilute sulphuric acid, precipitating TlCl from this solution; converting the TlCl into sulphate and precipitating the thallium from its sulphate solution by means of zinc.

Properties. Resembles lead in its appearance and physical properties. Specific gravity 11.8; melts at 290°.

COMPOUNDS OF THALLIUM.

Thallium acts both as monad and triad. As monad it forms a series of salts many of which are closely analogous to the salts of potassium. The hydroxide and carbonate are soluble in water with strong alkaline reaction (thallium is the only heavy metal which forms a soluble carbonate); the sulphate may replace potassium sulphate in alum.

All thallium salts are poisonous; some of them are occasionally used in making optical glass. Thallium compounds color the non-luminous gas flame a beautiful green.

Thallium Oxides. *Thallious Oxide,* Tl_2O, is a black powder, melting at 300°. Easily soluble in water with the formation of the *hydroxide,* TlOH, which is also formed by the action of water on the metal in the presence of air. *Thallic Oxide,* Tl_2O_3, formed by burning thallium in oxygen, yields Tl_2O at a red heat.

Thallium Chlorides. *Thallious Chloride,* TlCl, is formed by burning thallium in chlorine, or as a white precipitate on adding hydrochloric acid to a thallium solution. It is difficultly soluble in water and readily fusible. With platinic chloride it forms an almost insoluble double salt, $PtCl_4,2TlCl$.

Thallic Chloride, $TlCl_3$, is obtained when thallium chloride is treated with chlorine under water. Readily soluble in water; decomposes at 100° into TlCl and Cl_2.

Thallium Sulphates. *Thallious Sulphate,* Tl_2SO_4, is moderately soluble in water; forms double salts with magnesium and ferrous sulphates, and also with aluminium and chromium sulphates (thallium alums). *Thallic sulphate,* $Tl_23SO_4+7H_2O$, is decomposed by water with separation of the hydrated trioxide.

Thallious Sulphide, Tl_2S, is precipitated from an alkaline or acetic acid solution of a thallium salt by hydrogen sulphide as a black powder readily soluble in inorganic acids.

METALS OF THE COPPER GROUP.

COPPER. SILVER. MERCURY.

COPPER. Cu^{II} (Cuprum)=63.

Occurs abundantly and widely distributed. Native in considerable quantities. Combined as oxide, sulphide (often with other metallic sulphides) carbonate, etc

Extraction. Chiefly from the sulphide ores which contain much iron sulphide.

1. The ore is first roasted and then fused with a siliceous flux. In this way the iron is mostly oxidized and removed as slag, while a purer copper sulphide is left. This process is repeated until a nearly pure cuprous sulphide (Cu_2S) is obtained. This "white metal" is then roasted, the oxide formed reacting with unoxidized sulphide :—$Cu_2S+2Cu_2O=SO_2+4Cu$. The copper thus reduced is refined by fusion with coal, the foreign metals being removed by oxidation, and the small amount of copper oxide finally reduced by "poling."

2. From ores too poor to be profitably smelted, copper is obtained by the "wet process." The copper is first converted into sulphate by weathering or roasting, or into chloride by roasting with common salt; the sulphate or chloride is then brought into solution and the copper precipitated by iron.

Chemically pure copper is obtained by reduction of pure oxide by hydrogen ; or by electrolysis of a solution of pure sulphate.

Properties. Red metal, rather hard, exceedingly tough, malleable and ductile. Specific gravity 8.9. Melts higher than silver ; fused, it has a greenish-blue color and absorbs gases which escape through the crust as it cools. Heated in air it becomes coated with an easily detached black oxide. At ordinary temperatures it remains unaltered in dry air, but in moist air a green basic carbonate is formed on its surface. In dilute sulphuric and hydrochloric acids entirely insoluble except by the oxidizing aid of air, or in contact with a less positive metal (e. g. platinum). Dissolves in hot concentrated sulphuric acid with evolution of SO_2 ; readily soluble in nitric acid.

Used for boilers, stills, etc. ; for covering ships ; cartridge-cases ; as wire for electrical purposes ; as a source of copper preparations ; but most extensively in various alloys, especially in brass and bronze.

COMPOUNDS OF COPPER.

Copper forms two series of compounds : cuprous compounds in which a double atom of copper acts as a dyad (—Cu—Cu—), and cupric compounds in which the single dyad atom appears.

CUPROUS COMPOUNDS.

Cuprous Oxide, Cu_2O, occurs as cuprite and can be readily prepared by warming a solution of copper sulphate, caustic potash and grape sugar. A red crystalline powder ; dissolves in ammonia to a colorless solution which soon becomes blue from oxidation ; in hydrochloric acid to cuprous chloride ; in sulphuric and other oxygen acids to cupric salts with separation of metallic copper. Used to stain glass red.

Cuprous Chloride, Cu_2Cl_2, is readily prepared by leading SO_2 into a concentrated solution of copper sulphate and common salt. White crystalline powder, insoluble in water ; soluble in concentrated hydrochloric acid, in many chlorides and in ammonia. Its solutions absorb CO and certain gaseous hydrocarbons, especially acetylene, forming with them solid compounds.

Cuprous Sulphide, Cu_2S, occurs as copper glance. Formed by burning copper in sulphur vapor ; and by heating CuS in a current of hydrogen.

CUPRIC COMPOUNDS.

The anhydrous cupric salts are yellow or colorless ; the hydrated salts green or blue. Normal salts are generally soluble in water. Basic salts, which are readily formed, are usually insoluble and several of them serve as valuable pigments.

Copper Oxide, CuO, occurs in small quantities. Is prepared by heating copper in the air or oxygen ; or more easily by heating copper nitrate. Brownish black powder ; melts at a high temperature. At a red heat readily reduced by hydrogen, carbon and organic substances, and is hence used in organic analysis.

Copper Hydroxide, $Cu(OH)_2$ is obtained from copper solutions by caustic potash or soda as a voluminous blue precipitate, which is readily converted by heat, even when suspended in water, into black CuO. Readily soluble in ammonia forming a dark blue solution.

Copper Chloride, $CuCl_2$, dissolves readily in water, forming a green solution, from which it may be crystallized as $CuCl_2+2H_2O$. Deprived of its water of crystallization, it is a yellowish-brown deliquescent powder. At a red heat it gives cuprous chloride and chlorine. Soluble in alcohol, the solution burning with a fine green flame. Used in dyeing and cotton printing; in the preparation of copper pigments; to give "green fire" in pyrotechny, and as disinfectant.

Copper Sulphate, or *Blue Vitriol,* $CuSO_4$, occurs as the result of natural oxidation of copper sulphide, in the drainage water of copper mines. Obtained by roasting copper sulphide, and purified by crystallization. Crystallizes as $CuSO_4+5H_2O$ in transparent blue triclinic crystals, which are very soluble in water. Effloresces slowly in the air, loses 4 mol. of its crystallization water at 100°, forming a bluish-white powder, $CuSO_4+H_2O$, which becomes anhydrous and white above 200°. The anhydrous salt is very hygroscopic, becoming blue as it absorbs water. Used largely in dyeing and calico printing, in making copper pigments and other preparations; in electrotyping and electrical batteries and for many other purposes.

Copper Nitrate, $Cu2NO_3$, is very deliquescent and soluble in water and alcohol. Blue crystals, $Cu2NO_3+3H_2O$; or at a low temperature, $Cu2NO_3+6H_2O$. Is a strong oxidizing agent and used as such in dyeing and calico-printing.

Copper Carbonates. Basic carbonates alone are known. The carbonate $(HO)_2Cu_2CO_3$ occurs as malachite and is formed by the action of moist air on copper. Sodium carbonate precipitates from a cold solution of copper sulphate, $(HO)_2Cu_2CO_3+H_2O$, which on standing is converted into a green powder of the same composition as malachite. Malachite is used for vases and ornamental articles; finely powdered as green pigment. The carbonate, $(HO)_2Cu_32CO_3$, occurs as azurite and is used as blue pigment.

Copper Sulphide, CuS, occurs as indigo-copper and is obtained by leading H_2S into a solution of a cupric salt, as a brownish-black precipitate, which oxidizes readily in moist air to sulphate. Heated in a current of hydrogen it becomes Cu_2S.

Copper-Ammonium Compounds. Copper salts unite with gaseous or aqueous ammonia to form peculiar compounds which may be regarded as ammonium salts in which part of the hydrogen of the NH_4 group has been replaced by copper. E. g. $(NH_3)_2CuSO_4$.

DETECTION AND DETERMINATION OF COPPER.

Copper compounds color the non-luminous flame green or blue and give a borax bead which is green (hot) or blue (cold) in the oxidizing flame and red and opaque in the reducing flame. Copper is reduced from its solutions by iron or zinc and is readily detected by these and other reactions. It is weighed as metal, oxide or cuprous sulphide.

SILVER. Ag^I (Argentum)=107.7.

Occurs in all parts of the world. Free (native), occasionally in large masses, usually alloyed with copper and gold and sometimes containing mercury and other metals. More frequently combined: with sulphur; with sulphur and other metals; with chlorine and sometimes with bromine and with iodine. It is also found as carbonate, and is often contained in small quantities in argentiferous galena, which serves as one of the most important sources of silver.

Extracted from its ores by different methods, the choice of which depends upon the nature of the ore and economic considerations. a. A silver-lead alloy is formed from which the lead is then removed by oxidation (cupellation). b. The silver is dissolved in mercury, which is removed by distillation. c. The silver is brought into solution and precipitated by copper.

Properties. Silver is a white metal and takes a high polish. In a finely divided state, as obtained by reduction of the chloride and other salts, it is a dull grey powder. Specific gravity 10.57. Melts at 1040° and can be distilled at the heat of the oxy-hydrogen blow-pipe. When melted it absorbs twenty-two times its volnme of oxygen from the air (without being oxidized), which escapes through the exterior crust as it solidifies ("spitting" of silver). It is soft, tough, very malleable and ductile. Unaltered in pure air, but tarnishes in air containing suphuretted hydrogen, from the formation of a film of silver sulphide. Is slightly oxidized by ozone. Unites directly with the halogens and with sulphur. Insoluble in hydrochloric acid (which produces an insoluble superficial film of AgCl) and in aqua regia (which converts it into AgCl); dissolves readily in nitric acid, and slowly in concentrated boiling sulphuric acid with evolution of SO_2.

Uses. Pure silver is rarely used except in the chemical laboratory and for preparing pure salts. The silver used for plate, coin or ornamental purposes is alloyed with copper, which renders it harder and more sonorous.

COMPOUNDS OF SILVER.

The soluble salts (nitrate and sulphate) have a neutral reaction, harsh and metallic taste and are poisonous. Most of its salts are insoluble in water and are prepared from the nitrate by precipitation. They are all soluble in ammonia. Silver is precipitated from its solutions by many metals, phosphorus, phosphorous acid, sulphurous acid, iron sulphate and many organic substances.

Silver Oxide, Ag_2O, is precipitated from a silver nitrate solution by caustic alkalies or barium hydroxide. It is a brown powder, slightly soluble in water; decomposed into silver and oxygen at 250°; is reduced by hydrogen at 100°. Ignites readily oxidizable substances—e. g. Sb_2S_3, red phosphorus, etc.—when rubbed with them in a dry state. The hydroxide AgOH is unknown. If freshly precipitated Ag_2O is digested with ammonia, a black powder is obtained which explodes violently on the lightest touch, even when moist (Berthollet's fulminating silver).

Silver Chloride, AgCl, occurs as horn-silver and falls as a curdy white precipitate on adding hydrochloric acid or a soluble chloride to a silver solution. Melts at 260° and forms on cooling a horn-like mass. Insoluble in water and acids ; readily soluble in ammonia, potassium cyanide and sodium thiosulphate. Precipitated silver chloride when exposed to light turns violet and then black. Easily reduced to metallic silver by nascent hydrogen, many organic substances and by fusion with sodium carbonate. Used in photography and for the preparation of pure silver.

Silver Bromide, AgBr, and **Silver Iodide**, AgI, occur as rare minerals. They resemble silver chloride and are used in photography.

Silver Sulphate, Ag_2SO_4, is sparingly soluble in water ; more soluble in dilute sulphuric or nitric acid.

Silver Nitrate, $AgNO_3$, is prepared by dissolving ordinary silver (containing copper) in nitric acid, evaporating to dryness and heating until the copper nitrate is completely decomposed. The silver nitrate is then dissolved, filtered from the copper oxide and crystallized. Readily soluble in water, crystallizes in large rhombic plates. Melts at 198° and at a red heat yields oxygen, NO_2 and silver. Unaltered in the air and in light; is blackened (reduced) by organic substances. Is corrosive and very poisonous.

Silver nitrate is the most important salt of silver. It is often cast into sticks and is used in this form as a powerful cautery (luna caustic). Employed largely in photography; used also in making indelible inks, in medicine and as a reagent in the laboratory.

Silver Sulphide, Ag_2S, occurs as argentite and is formed by heating silver with sulphur or precipitating a silver solution with H_2S. Brownish black powder, soluble in hot nitric acid; converted by careful heating into Ag_2SO_4.

Silver-Ammonium Compounds. Silver salts form compounds with ammonia, which resemble the copper-ammonium compounds.

DETECTION AND DETERMINATION OF SILVER.

Silver is readily detected by the precipitation of its chloride, which is soluble in ammonia, and is almost always weighed as chloride.

MERCURY. Hg^{II} (Hydrargyrum)=200.

Occurs native to a small extent. In combination it is found chiefly as sulphide; rarely as chloride, and silver amalgam.

Extraction. From the sulphide (cinnabar) (a) by distillation in a current of air:—$HgS+O_2=SO_2+Hg$; or (b) by distillation with lime:—$4HgS+4CaO=3CaS+CaSO_4+4Hg$.

Commercial mercury may be purified from the foreign metals it contains by distillation or treatment with oxidizing agents (e. g. dilute nitric acid, through a long column of which it is allowed to flow in a very thin stream).

Properties. Mercury is the only metal liquid at ordinary temperatures. Specific gravity 13.6. It solidifies at —39°.5 and in the solid state is soft, malleable and ductile. It crystallizes in regular octahedrons. Volatilizes slightly at ordinary temperatures and boils at 360°. Unaltered in air at ordinary temperatures ; near its boiling point it oxidizes slowly to HgO. Insoluble in dilute hydrochloric and sulphuric acids ; dissolves in hot concentrated sulphuric acid with evolution of SO_2 ; readily soluble in nitric acid. Mercury dissolves many metals forming amalgams.

Uses. Mercury is used in barometers, thermometers and other physical instruments and for many purposes in physical and chemical laboratories ; in the extraction of gold and silver ; for making many useful amalgams ; for the preparation of mercury compounds, and in a finely divided state in medicine.

COMPOUNDS OF MERCURY.

Mercury forms two series of salts. In the *mercurous* compounds a double atom $Hg_2{}^{II}=(—Hg—Hg—)$ acts as dyad ; while the *mercuric* compounds contain a single atom with full valence (Hg^{II}). The mercurous compounds are analogous to the cuprous and silver compounds. The mercuric compounds resemble the cupric and also those of zinc and cadmium. The mercury salts are generally formed by the action of an acid upon mercury: the mercuric, when the acid is hot, or in excess; the mercurous, when the mercury is in excess and the acid cold. A salt of one series is readily converted into the corresponding salt of the other series by oxidation or reduction. Many mercurial preparations are of therapeutic importance.

MERCUROUS COMPOUNDS.

Mercurous Oxide, Hg_2O, precipitated by caustic soda from mercurous solutions as a black powder, which decomposes in the light or at 100° into mercury and HgO. The hydroxide is unknown.

Mercurous Chloride, or *Calomel,* Hg_2Cl_2, occurs as horn-quicksilver and is precipitated by hydrochloric acid from a mercurous solution. Usually prepared by sublimation from a mixture of mercury and $HgCl_2$; or mercury, $HgSO_4$ and NaCl. Mercurous chloride, on heating, sublimes without fusing. Insoluble in water, alcohol and dilute acids. Moistened with ammonia it blackens, being converted into mercurous-ammonium chloride, NH_2Hg_2Cl. Used largely in medicine.

Mercurous Nitrate, Hg_22NO_3, formed by the action of cold dilute nitric acid on mercury. Crystallizes as $Hg_22NO_3+2H_2O$. Dissolves unchanged in a small quantity of water; decomposed by a larger quantity into an acid salt which dissolves, and an insoluble basic salt, $(HO)Hg_2NO_3$. Heated, it yields HgO. In nitric acid solution it slowly oxidizes to mercuric nitrate; this is prevented by the presence of mercury.

MERCURIC COMPOUNDS.

Mercuric Oxide. *Red Oxide of Mercury,* HgO, is formed slowly when mercury is heated in air to a temperature a little below its boiling point, and is precipitated from mercuric salts by caustic soda; usually prepared by heating a mixture of mercury and mercuric nitrate. Formed by precipitation, it is a yellow amorphous powder; prepared in the dry way it is a red crystalline powder. Slightly soluble in water. Heated it becomes black and decomposes into mercury and oxygen at 400°.

Mercuric Chloride, or *Corrosive Sublimate,* $HgCl_2$, is usually prepared by sublimation from a mixture of common salt and mercuric sulphate:—$HgSO_4+2NaCl=Na_2SO_4+HgCl_2$. Mercuric chloride is soluble in water and crystallizes in rhombic prisms. More soluble in alcohol and ether than in water. Melts at 293°, boils at 303°, volatilizing more readily than Hg_2Cl_2. Readily reduced to Hg_2Cl_2. Coagulates albumen and is very poisonous. Used as a preservative for wood; as vermin poison; in medicine and in making other mercurial preparations.

Mercuric Sulphate, $HgSO_4$, prepared by dissolving mercury in hot concentrated sulphuric acid. Sparingly soluble in water and decomposed by water into an insoluble basic salt. At a red heat yields mercury, SO_2 and oxygen.

Mercuric Nitrate, $Hg2NO_3$, is formed by heating mercury with an excess of nitric acid; but is obtained pure with difficulty as it is very deliquescent and is readily decomposed by water into insoluble basic salts.

Mercuric Sulphide, HgS, occurs as cinnabar. Formed by rubbing together mercury and flowers of sulphur and by precipitating a mercuric solution with hydrogen sulphide. The artificially prepared sulphide is a black powder, insoluble and unchanged by dilute acids ; soluble in concentrated nitric acid or in aqua regia. Heated, out of contact with air, it sublimes, forming a dark red crystalline mass resembling cinnabar.

Used in the black amorphous form in medicine ; in the red modification as pigment (vermillion), often adulterated with red lead or red oxide of iron.

Mercury-Ammonium Compounds analogous to those of copper and silver are formed by the action of dry and aqueous ammonia on mercury salts.

DETECTION AND DETERMINATION OF MERCURY.

Mercury is readily detected by its wet reactions and by the reduction of its compounds on heating in a dry tube with sodium carbonate. It is weighed as mercury, mercuric sulphide, or mercurous chloride.

METALS OF THE CERIUM GROUP.

Yttrium. $Y=89.6$.
Lanthanum. $La=138.5$
Cerium. $Ce=141.2$.
Didymium. $Di=145$.
Erbium. $Er=166$.
Ytterbium. $Yb=172.6$.

These metals and probably several others (terbium, scandium, samarium, etc.) occur chiefly as silicates in several rare minerals in Scandinavia, Siberia and Greenland. The separation of the metals of this group and the preparation of their pure compounds is extremely difficult. Cerium, lanthanum, didymium and yttrium (impure) have been obtained in the metallic state. Cerium oxalate is used in medicine.

METALS OF THE ALUMINIUM GROUP.

ALUMINIUM. INDIUM. GALLIUM.

ALUMINIUM. $Al^{IV}=27$.

Occurrence. Next to oxygen and silicon the most abundant and widely distributed element. Occurs only in combination ;—as oxide, hydroxide, fluoride and phosphate, but chiefly as silicate, both by itself and as constituent of very many minerals and rocks.

Prepared from its chloride (usually the double chloride of sodium and aluminium is used), or from cryolite (double fluoride of aluminium and sodium) by reduction with sodium.

Properties. White metal, about as hard as silver, malleable, ductile and sonorous. Specific gravity 2.6. Melts about 600° and can not be volatilized. Does not tarnish in air and oxidizes but slightly at a red heat. Dissolves readily in hydrochloric acid and caustic alkali ; slowly in warm sulphuric acid ; is not attacked by nitric acid. Soluble in organic acids in the presence of common salt and hence is not adapted for domestic utensils.

Used for making opera glasses, weights for balances and certain physical instruments, and in certain alloys. Its wider application is prevented by the cost of its preparation.

COMPOUNDS OF ALUMINIUM.

Aluminium Oxide, or *Alumina,* Al_2O_3, occurs as corundum (sapphire and emery). Amorphous alumina is obtained by heating the hydroxide. In the heat of the oxy-hydrogen blowpipe it fuses to a thin liquid which crystallizes on cooling. In crystalline state it stands next to the diamond and boron in hardness. The natural oxide and the artificial oxide after being strongly heated are almost entirely insoluble in acids. In the form of emery it is used for grinding and polishing gems, steel, etc.

Aluminium Hydroxides occur native in certain minerals. $Al_2(OH)_6$ is formed as a gelatinous precipitate on adding caustic alkali to an aluminium solution. On heating it is converted into a tough horn-like mass and finally into Al_2O_3. The precipitated hydroxide has the property of withdrawing from solutions both inorganic and organic substances, and is, in consequence, used to purify water, clarify syrup, etc., and also as an important mordant in dyeing and cotton-printing. For this last purpose it is precipitated upon the fibres of the cloth from some soluble salt of aluminium. The coloring matters precipitated by means of aluminium compounds are also employed as pigments (lakes) for coloring carpets and paper.

Aluminates. Aluminium hydroxide acts toward strong bases as a weak acid, forming with them salts called aluminates.

Sodium Aluminate, $Al_2(ONa)_6$, is formed by dissolving the hydroxide in caustic soda. It is prepared on a large scale by heating the natural hydroxide (bauxite) with carbonate of soda ; also by heating cryolite with chalk (cf. p. 24). Soluble in water; converted by carbon dioxide into aluminium hydroxide. Used as mordant; in the preparation of lake colors ; for making artificial stone, weighting paper, etc.

Calcium Aluminates are insoluble in water, and most of them are readily fusible, so that lime and alumina act as fluxes toward each other.

Magnesium Aluminate, $MgAl_2O_4$, occurs as spinelle.

Beryllium Aluminate, $BeAl_2O_4$, as chrysoberyl.

Aluminium Chloride, Al_2Cl_6, crystallizes out of a concentrated solution of alumina in hydrochloric acid as $Al_2Cl_6+12\ H_2O$, and decomposes on heating into Al_2O_3 and HCl. The anhydrous chloride is prepared by heating an intimate mixture of alumina and carbon in a current of dry chlorine. It sublimes without fusing, forming a colorless, crystalline, very deliquescent mass.

It forms double salts with many metallic chlorides, one of which, $Al_2Cl_6,2NaCl$, prepared by sublimation from a mixture of alumina, common salt and carbon heated in a current of dry chlorine, is used for the preparation of aluminium, as it is much less hygroscopic than Al_2Cl_6.

Aluminium chloride is used in concentrated solution as antiseptic and disinfectant (chloralum).

Aluminium Sulphate, Al_23SO_4, is prepared on a large scale by the action of sulphuric acid on china-clay. It crystallizes with difficulty with 18 mol. of

H_2O. Occurs in this form as feather alum. Readily soluble; dissolves alumina forming basic salts; unites with alkali sulphates forming double sulphates called alums.

Used in an impure state under the name of concentrated alum in place of alum, especially in dyeing and the manufacture of paper.

THE ALUMS.

An alum is a double sulphate of aluminium, (iron chromium or manganese) and an alkali metal (or thallium). All alums crystallize in octahedrons or cubes and have the general formula :—$AlX2SO_4+12H_2O$, or $Al_23SO_4,X_2SO_4 + 24H_2O$. Ordinary alum is

Potassium Alum, $Al_23SO_4,K_2SO_4+24H_2O$. An insoluble basic alum, $K_2SO_4,Al_23SO_4,2Al_2(OH)_6$, occurs as alum-stone in volcanic regions. Commercial alum is obtained (a) from alum-stone by roasting the alum-rock, lixiviating and crystallizing (Roman alum); (b) in large quantities by adding potassium salts (usually the chloride) to aluminium sulphate obtained by the action of sulphuric acid on various natural compounds of aluminium.

It crystallizes in large octahedrons. Very much more soluble in hot than in cold water. Insoluble in alcohol. Effloresces in the air and loses 18 mol. of water in perfectly dry air. Melts at 92° in its water of crystallization; heated higher, it loses water, swelling up into a porous anhydrous mass of "burnt alum," which dissolves very slowly in water. Is decomposed by stronger heat, being finally converted into potassium aluminate. Used as an important mordant and for a variety of other technical purposes, as disinfectant and as an astringent in medicine.

Aluminium Silicates occur in combination with other silicates in great abundance, forming the chief constituents of the earth's crust.

The clays employed for making porcelain, earthenware and bricks are more or less impure hydrated silicates of aluminium, which form with water a plastic mass and when heated bake together forming a hard but porous solid which does not soften in water.

Ultramarine is a double silicate of aluminium and sodium containing sulphur. It occurs as lapis lazuli and is made by heating a mixture of clay with sulphate (or other compound) of sodium, sulphur, carbon and sand.

Aluminium Sulphide, Al_2S_3, is decomposed by water into $Al_2(OH)_6$ and H_2S ; hence it is not formed in the wet way.

DETECTION AND DETERMINATION OF ALUMINIUM.

Aluminium is readily detected by the behavior of its hydroxide. It is determined as oxide, by ignition of the precipitated hydroxide.

INDIUM. In=113.4. GALLIUM. Ga=69.9.

Both these rare metals occur in certain zinc blendes. They were discovered in 1863 and 1875 respectively by means of spectrum analysis.

Indium is prepared from its oxide by ignition in a current of hydrogen or by fusion with sodium. It is a white, soft, malleable metal, melting at 176°. Its compounds color the non-luminous flame dark blue.

Gallium is prepared by electrolysis of an alkaline solution. It has a bluish-white color and fuses at 30°.1. Its spectrum consists of two violet lines. The compounds of both metals are analogous to those of aluminium.

METALS OF THE IRON GROUP.

IRON. MANGANESE. COBALT. NICKEL.

IRON. $Fe^{II \text{ and } IV} = 56$.

Occurs rarely native in meteorites, in minute quantities in certain basalts and lavas, and in considerable masses (natural steel) formed by the firing of coal-pits. In combination it is very abundant and widely distributed, occurring chiefly as sulphide, oxide, hydroxide and carbonate. Compounds of iron are also found in most minerals, rocks, soils and natural waters (chalybeate springs) and in plants (a necessary constituent of chlorophyll) and animals (especially in the blood).

CAST-IRON.

Extraction. Iron is extracted from the natural oxides or from those produced by the roasting of other ores (chiefly hydroxides and carbonates) by reduction with carbon. The process is almost always carried out in "blast furnaces" and is a continuous one during the life of the furnace. Charges of roasted ore, flux (limestone or clay according to the nature of the rock in which the ore occurs) and fuel (charcoal, coal or coke) in the proper proportions, are introduced at the top of the tall cylindrical furnace, while at the bottom the slag flows out continuously and the molten iron is withdrawn at regular intervals and run into sand moulds forming semi-cylindrical bars or "pigs."

Properties. The crude pig- or cast-iron obtained as the product of the blast furnace, is brittle, granular and far from pure, containing 2.5–5.9 per cent. of carbon, and 8–11 per cent. of other elements, chiefly phosphorus, silicon, sulphur and manganese. It may be divided into two chief classes, grey cast-iron and white cast-iron, according to the state of the carbon which it contains.

a. *Grey cast-iron* contains its carbon chiefly in the form of graphite. It is formed when molten iron saturated with carbon at a high temperature is allowed to cool slowly. Specific gravity 7.0–7.4; melts at 1500–1700° and is well adapted for castings, as it expands on solidification, and is soft enough to be planed, filed, bored, etc.

b. *White cast-iron* contains the whole of its carbon in a state of chemical combination. It is formed when the temperature of the furnace is not high enough for complete carburisation of the iron; or by quickly cooling melted grey iron. It is very hard and brittle. Specific gravity 7.4–7.6; melts at 1400–1500°, but is less liquid than grey-iron and contracts considerably on solidification and is thus unfit for casting. It is used chiefly for the preparation of wrought-iron and steel. When allowed to cool slowly from fusion white-iron is converted into grey-iron.

Spiegel or specular iron is a white-iron containing the largest proportion of carbon, and manganese as an essential constituent. It is extremely hard and highly crystalline. Specific gravity 7.6–7.7 ; melts at 1500–1600° ; used chiefly in the manufacture of Bessemer steel.

c. *Mottled-iron* is intermediate between grey and white-iron. It has a granular structure ; gives harder and stronger castings than grey-iron and can be used both for foundry purposes and for making wrought-iron and Bessemer steel.

WROUGHT–IRON.

Wrought-iron is made (a) chiefly from white cast-iron by removing most of the carbon and other substances by oxidation (puddling process). (b) Directly from ores of good quality by reduction in bloomery forges.

Properties. Wrought-iron contains 0.1–0.7 per cent. carbon and is the purest commercial form of iron. When the carbon is not more than 0.3 per cent., the texture is fibrous; when more than this, granular or crystalline. The hardness increases with the amount of carbon. The presence of sulphur renders it brittle while hot ("red-short") ; phosphorus makes it brittle while cold ("cold-short"). Its specific gravity is 7.6–7.9 ; it melts 1900–2100°. It is very tenacious, soft, malleable and can be welded.

STEEL.

Steel is produced (a) by carburizing wrought-iron. Bars of iron are packed in powdered charcoal or soot and kept at a full red heat for several days, the time depending upon the quality of the steel desired (cementation or blister-steel). *Cast-steel* is obtained from blister-steel by fusion in plumbago crucibles.

(b) *Bessemer steel* is manufactured in large quantities from white cast-iron by burning out the carbon, silicon and other substances and then adding enough spiegel-iron to give the requisite amount of carbon. In from twenty to thirty minutes, five to twelve tons of cast-iron are converted into nearly the same weight of cast-steel.

(c) Steel is also produced by dissolving malleable scrap-iron in molten pig-iron (Siemens-Martin process).

Properties. Steel contains 0.75–2.4 per cent. of carbon, chiefly chemically combined. Specific gravity 7.62–7.92. It unites the advantages of cast-iron and wrought-iron, being fusible (1700–1900°), and malleable. It can also be welded not only on steel but also on iron. When quickly cooled from a red heat, steel becomes extremely hard and brittle, and almost perfectly elastic. The hardness and brittleness may be partially removed by heating it to 230-315°, and then allowing it to cool (tempering).

Case-hardening is effected by converting the surface of soft-iron into steel by heating it with cast-iron turnings or powdered charcoal.

Malleable castings are made by converting cast-iron superficially into wrought-iron or steel by heating it with various oxides.

PURE IRON.

Pure iron may be prepared by reduction of its oxide or chloride in a current of hydrogen or by electrolysis of the sulphate.

Properties. Pure iron has a specific gravity of 7.84. It is the most tenacious of all the ductile metals except nickel and cobalt; it becomes soft at a red-heat; at a white-heat can be welded; above the welding point is brittle; is more difficultly fusible than wrought-iron. All varieties of iron and steel are attracted by a magnet and become magnetized—steel permanently so.

Iron reduced by hydrogen at the lowest possible temperature forms a black powder which oxidizes at once with incandescence on exposure to the air; if reduced at a high temperature this does not occur.

Iron is unaltered in dry air, and in water free from air; but in moist air or in contact with water and air it becomes coated with hydroxide (rusts). Iron decomposes steam at a red heat. All varieties of iron dissolve readily in most dilute acids with evolution of hydrogen and hydrocarbons from the combined carbon, and separation of black scales of graphite when this is contained in the iron. Dipped into concentrated nitric acid, iron becomes covered with a film of oxide and is rendered "passive." In contact with zinc, iron becomes electro-negative and is thus protected from rusting.

Uses. Iron is the most important and useful of the metals.

COMPOUNDS OF IRON.

Iron forms two well defined series of compounds. In the *ferrous* compounds it acts as a dyad. The compounds of this series are analogous to those of the magnesium group. In the *ferric* compounds iron acts as a tetrad, usually in a hexad double atom $(Fe{-}Fe)^{VI}$, forming compounds analogous to those of aluminium and chromium. Iron salts are of great importance in the arts and in medicine. In the arts they are chiefly employed as mordants; in medicine as tonics and astringents.

FERROUS COMPOUNDS.

The anhydrous ferrous salts of colorless acids are usually colorless; the hydrated salts are greenish-blue. The soluble salts have a sweet, astringent taste; their solutions absorb oxygen readily, forming insoluble basic ferric salts when no free acid is present. They are consequently powerful reducing agents and are largely used as such in the laboratory and in the arts.

Solutions of ferrous salts absorb nitric oxide forming a dark solution, which is decomposed by heat.

Ferrous Oxide, FeO, is formed by heating ferrous oxalate or by heating ferric oxide in a current of CO and CO_2. Black powder, which oxidizes very readily to Fe_2O_3.

Ferrous Hydroxide, $Fe(OH)_2$, is formed on adding caustic soda to a ferrous solution in absence of air, as a white precipitate which rapidly absorbs oxygen from the air, becoming a dirty green color and finally brown $Fe_2(OH)_6$ It is very slightly soluble in water.

Ferrous Chloride, $FeCl_2$, is prepared in an anhydrous state by heating iron in a current of chlorine or hydrochloric acid gas. It forms colorless shining scales which melt at a red heat and sublime at a higher temperature. Exposed to air it decomposes into Fe_2Cl_6 which volatilizes, and Fe_2O_3. It is deliquescent and very soluble in water from which it crystallizes in bluish-green monoclinic prisms as $FeCl_2+4H_2O$. The hydrated salt is also obtained by evaporating a solution of iron in hydrochloric acid. The crystals are deliquescent and oxidize in the air.

Ferrous chloride in solution is used in dyeing silk and wool blue (with ferricyanide of potassium), and in medicine.

Ferrous Iodide, FeI_2, is prepared by direct union of iron and iodine. Grey crystalline mass which resembles the chloride in its properties. Used in medicine.

Ferrous Sulphate, $FeSO_4$, is formed by dissolving iron in sulphuric acid and by the natural oxidation of iron pyrites. It is obtained in large quantities from the latter source as a bye-product in several industries. Crystallizes as $FeSO_4+7H_2O$ in monoclinic prisms. When pure the salt has a pale blue color; usually, however, it is green from the presence of a small amount of ferric (yellow) compounds. Heated, it loses six mol. of water at 100° and yields a white powder; at about 300° the seventh mol. of water is driven off; at a higher temperature the salt is decomposed into Fe_2O_3, SO_3 and SO_2. Readily soluble in water. Both the crystals and their solution oxidize slowly in the air, being gradually changed into brownish insoluble basic ferric sulphate. The solution absorbs nitric oxide, becoming dark brown. This reaction is used in the laboratory for the detection of nitric acid.

Ferrous sulphate forms double salts with ammonium and potassium sulphates, which are far more stable than the ferrous sulphate alone. E. g. $(NH_4)_2SO_4.FeSO_4+6H_2O$.

Ferrous sulphate is the most important of the salts of iron. It is largely used under the name of "copperas" or "green vitriol" as a mordant; in making ink and Prussian blue; as a disinfectant; as a precipitant of gold and for many other purposes.

Ferrous Carbonate, $FeCO_3$, occurs as spathic iron forming an important iron-ore. Formed on adding alkali carbonate to a ferrous solution as a white precipitate, which rapidly becomes discolored, decomposing finally into $Fe_2(OH)_6$. Soluble in water containing carbon dioxide (cf. calcium carbonate) and is hence found in chalybeate springs.

Ferrous Sulphide, FeS, is readily formed by the direct union of the elements, when iron filings and sulphur are melted together or rubbed up with water, and is precipitated as a black powder on adding ammonium sulphide to an iron solution. If the solution is that of a ferric salt, the sulphide is mixed with sulphur. The sulphide, as usually prepared in the dry way, is a grey metallic mass which dissolves readily in acids with evolution of H_2S, and is used for the preparation of this gas.

FERRIC COMPOUNDS.

The ferric salts are formed by dissolving ferric oxide in acids or by oxidation of ferrous salts in the presence of free acid.

The anhydrous ferric salts are usually colorless; the hydrated, yellow or brown. The soluble salts have a peculiar astringent taste and are readily reduced to the corresponding ferrous compounds. They crystallize with difficulty.

Ferric Oxide, Fe_2O_3, is one of the most important ores of iron, occurring as hæmatite or specular iron ore, in many varieties. It is formed in general by igniting any iron compound containing oxygen in the air. Is obtained in large quantities as bye-product in the preparation of fuming sulphuric acid from green vitriol, forming "colcothar" or "caput mortuum." The ignited as well as the natural ferric oxide is very difficultly soluble in acids. It is readily reduced at a red heat by hydrogen, carbon or carbon monoxide.

The artificial oxide is used as rouge for polishing glass and metals and as the chief constituent of a number of paints.

Ferric Hydroxide, $Fe_2(OH)_6$, is formed as a reddish-brown precipitate on adding ammonia to a ferric solution. On heating at 100° or boiling with water, the hydroxide becomes darker and denser, changing into the basic hydroxide $Fe_2O_3,Fe_2(OH)_6$. This basic hydroxide occurs naturally as brown hæmatite, limonite and ochre and is formed as iron rust. Ferric hydroxide acts as a carrier of oxygen from the air to oxidizable substances and hence serves as an antiseptic, favors the formation of nitre in the soil, corrodes wood (rusty nails) and cloth. Precipitated from solution upon cloth fibres, it is an efficient mordant. Freshly precipitated ferric hydroxide dissolves in a solution of ferric chloride and can be separated from the latter by dialysis as *soluble ferric hydroxide,* a blood-red solution which is used in medicine.

Ferroso-Ferric Oxide or *Magnetic Oxide of Iron,* Fe_3O_4, occurs as magnetite. It is formed by heating iron in a current of steam. The natural oxide is magnetic (lodestone or natural magnet).

Ferric Acid, $FeO_2(OH)_2$, is not known, but its potassium salt is formed by heating iron filings with saltpetre; ferric oxide with potassium hydroxide and saltpetre; or by leading chlorine through a solution of caustic potash in which freshly precipitated ferric hydroxide is suspended. On evaporation of its solution, potassium ferrate crystallizes in dark red prisms. The solution gradually decomposes with separation of ferric hydroxide and oxygen.

Ferric Chloride, Fe_2Cl_6, is formed in the anhydrous state by heating iron in a current of dry chlorine or amorphous ferric oxide in HCl. Black iridescent scales or dark green plates of metallic lustre, which volatilize more readily than ferrous chloride. Deliquescent and very soluble in water, alcohol and ether. Ferric chloride can also be obtained in solution by oxidizing a solution of ferrous chloride with chlorine, or with nitric acid in the presence of hydrochloric acid. On evaporation to syrupy consistence, the yellowish deliquescent hydrate, $Fe_2Cl_6+12H_2O$, is deposited. On heating, this decomposes into H_2O, HCl, Fe_2O_3 and anhydrous Fe_2Cl_6, which sublimes.

Ferric chloride is used as a mordant, as a styptic in medicine and as reagent in the laboratory.

Ferric Sulphate, Fe_23SO_4, is formed by the oxidation (by nitric acid) of ferrous sulphate in the presence of sulphuric acid. The yellowish-brown solution on evaporation yields the anhydrous salt as white powder. Ferric sulphate unites with the alkali sulphates to form "iron alums."

Ferric Nitrate, Fe_26NO_3. The brown concentrated solution deposits, on addition of nitric acid, colorless deliquescent crystals, which contain 12 or 18 mol. of water of crystallization. The solution is used as a mordant.

Ferric Sulphides. *Iron Pyrites,* FeS_2, occur abundantly, usually in very hard yellow cubical crystals of metallic lustre. Used largely as a source of sulphur dioxide in the manufacture of sulphuric acid, and also for the preparation of sulphur and green vitriol.

Magnetic Pyrites, composition varying between Fe_7S_8 and Fe_8S_9, occurs naturally. It is attracted by the magnet and is sometimes itself magnetic.

Potassium Ferrocyanide, *yellow prussiate of potash,* $K_4Fe(CN)_6$, is formed by digesting a solution of potassium cyanide with ferrous compounds or finely divided iron. Prepared on a large scale by heating crude potash with iron and nitrogenous organic matter (horn, leather-clippings, dried blood, etc). On lixiviating the mass thus obtained, the ferrocyanide is formed by the reaction of the potassium cyanide and iron sulphide produced in the first operation.

Potassium ferrocyanide is deposited from solution in large yellow crystals, $K_4Fe(CN)_6+3H_2O$, which lose their water of crystallization at 100°, falling into a white powder. Heated with dilute sulphuric acid, hydrocyanic acid is evolved ; with concentrated sulphuric acid, carbon monoxide. Added to a ferrous solution the salt gives a white precipitate ; in a ferric solution the precipitate is "Prussian blue."

Potassium Ferricyanide, $K_3Fe(CN)_6$, is prepared from the ferrocyanide by oxidation with chlorine. Large dark-red prisms, which give a yellowish-brown aqueous solution. It is a powerful oxidizing agent. Added to a ferrous solution it precipitates "Turnbull's blue ;" in a ferric solution no precipitate is produced.

DETECTION AND DETERMINATION OF IRON.

Ferrous and ferric salts are readily detected and distinguished by their reactions with ferro- and ferricyanide of potassium. Iron compounds color the borax bead green in the reducing, yellow in the oxidizing flame.

For gravimetric determination iron is precipitated from a ferric solution as hydroxide by ammonia and weighed as Fe_2O_3. Iron is determined volumetrically by the reduction of potassium permanganate by a ferrous salt.

MANGANESE. $Mn^{II \text{ and } IV} = 54.8$.

Occurs only in combination; quite abundantly as dioxide, MnO_2, and in smaller quantities as Mn_2O_3, Mn_3O_4, hydroxide, carbonate and sulphide. Its compounds are also contained in minute quantities in many other minerals, especially in silicates. and iron ores.

Prepared from its oxides by reduction with carbon at a white heat. Like iron, it takes up carbon and silicon, and, since its oxides are not reduced by hydrogen, is obtained pure with difficulty.

Properties. Grey or reddish-white metal, very hard and brittle. It is non-magnetic. Specific gravity 8.0. Difficultly fusible. Oxidizes quickly in the air; decomposes boiling water and dissolves very easily in dilute acids. "Cast manganese" containing iron and other impurities is unaltered in the air.

COMPOUNDS OF MANGANESE.

Like iron, manganese forms two series of compounds (the manganous and manganic) in which it acts as dyad and as tetrad (usually in a hexad double atom Mn_2^{VI}) respectively.

MANGANOUS COMPOUNDS.

The manganous salts are usually of a pale red color. They are more stable than the corresponding ferrous compounds.

Manganous Oxide, MnO, is formed by heating any of the other oxides or the carbonate in hydrogen. Greenish powder. On heating in the air it is converted into Mn_3O_4.

Manganous Hydroxide, $Mn(OH)_2$, is obtained on adding caustic alkali to a manganous solution as a white precipitate, which soon oxidizes and turns brown.

Manganous Chloride, $MnCl_2$. Obtained in the anhydrous state by heating the oxide or carbonate in HCl; pale rose-colored mass which melts at a red heat. It crystallizes from solution as $MnCl_2+4H_2O$. On heating, the crystals lose water and decompose partly into hydrochloric acid and oxide. Forms a double chloride with ammonium, from which the anhydrous $MnCl_2$ may be obtained by heating (cf. magnesium chloride). Obtained in large quantities as bye-product in the preparation of chlorine in the manufacture of bleaching powder. Used to some extent in dyeing and cotton-printing, and as disinfectant.

Manganous Sulphate, $MnSO_4$, crystallizes below 6° as $MnSO_4+7H_2O$, isomorphous with ferrous sulphate; between 7°–20° as $MnSO_4+5H_2O$, isomorphous with copper sulphate; between 20°–30° as $MnSO_4+4H_2O$. At 200° the salt becomes anhydrous. Used in dyeing and cotton-printing and is an excellent "drier" for paint.

Manganous Carbonate, $MnCO_3$, occurs as manganese spar; obtained artificially on adding alkali carbonate to a manganous solution, as a white precipitate which soon becomes brown. Soluble in water containing carbon dioxide (cf. calcium carbonate).

Manganous Sulphide, MnS, occurs as manganese blende. Artificially prepared by heating the carbonate or oxide in a current of hydrogen sulphide, it forms a dark grey powder. On adding an alkali sulphide to a manganous solution a flesh-colored precipitate of the hydrated sulphide is obtained, which becomes brown on exposure to the air.

MANGANIC COMPOUNDS.

The manganic salts are much less stable than the corresponding ferric salts.

Manganic Oxide, Mn_2O_3, occurs as braunite. Obtained as a black powder by igniting any oxygen compound of manganese in oxygen.

Manganic Hydroxide, $Mn_2O_2(OH)_2$, occurs as manganite and is formed by the oxidation of manganous hydroxide in moist air, as a dark brown powder.

Mangano-Manganic Oxide, Mn_3O_4, occurs as hausmanite. Is formed by heating any oxide of manganese in air. Brown powder. With hydrochloric acid it yields manganous chloride and chlorine; with concentrated sulphuric acid, manganous sulphate and oxygen.

Manganic Sulphate, Mn_23SO_4, is obtained by dissolving the hydrated peroxide in sulphuric acid. Dark green deliquescent powder which yields a violet solution. The solution is unstable, manganic hydroxide separating out on standing; warmed it decomposes into $MnSO_4$, H_2SO_4 and oxygen. With potassium and sodium sulphates it unites to form alums, which are somewhat more stable.

Manganese Dioxide, *Manganese Peroxide* or *Black Oxide of Manganese,* MnO_2, occurs as pyrolusite, the most abundant and most important compound of manganese. The dioxide may be artificially obtained (regeneration of chlorine residues) by heating manganous carbonate, or a mixture of manganous hydroxide with an excess of lime, in a current of hot air. It is also precipitated in the hydrated state from a manganous solution by sodium hypochlorite, or by leading chlorine into the solution after addition of sodium carbonate. Gently heated it is converted into Mn_2O_3, at a strong red heat into Mn_3O_4 with evolution of oxygen. Insoluble in nitric acid, dissolves in hydrochloric acid with evolution of chlorine; in concentrated sulphuric acid yielding oxygen.

The natural dioxide is used in immense quantities in making chlorine; and is also employed in the preparation of bromine and iodine; in the laboratory for making oxygen and as an oxidizing agent; in glass-making to produce colorless and violet glass ("glass-makers' soap"); as a constituent of matches; as a source of manganese compounds and for various other purposes.

Manganic Acid, H_2MnO_4, is known only in solution and its anhydride, MnO_3, is unknown.

Potassium Manganate, K_2MnO_4, is produced by the oxidation of manganese compounds in the presence of caustic potash, and may be prepared by fusion of a mixture of caustic potash and manganese dioxide with saltpetre or potassium chlorate. The dark green mass thus obtained gives an unstable green solution, from which crystals of the manganate may be obtained by evaporation in a vacuum. The solution decomposes on standing, more quickly on heating, into a red solution of permanganate, $KMnO_4$, with separation of hydrated dioxide.

Permanganic Acid, $HMnO_4$, is known only in unstable aqueous solutions. Its anhydride, Mn_2O_7, is obtained as a very unstable oily liquid, on adding potassium permanganate to concentrated sulphuric acid. The anhydride decomposes slowly on standing, explosively when heated, and is a very powerful oxidizing agent.

Potassium Permanganate, $KMnO_4$, is prepared by lixiviating the green mass of potassium manganate, obtained as above, with boiling water. Deposits in dark red crystals of metallic lustre which dissolve in 16 parts of cold water, forming a deep purple solution. Potassium permanganate is a very strong oxidizing agent, destroying organic substances even in dilute solution, and hence its solution cannot be filtered through paper. It sets chlorine free from hydrochloric acid, oxidizes sulphurous acid to sulphuric, ferrous salts to ferric, etc. In these oxidations its color is destroyed on account of reduction to manganous salts. Used as a disinfectant ; for bleaching, dyeing and as an important reagent in the laboratory, especially in volumetric analysis.

Sodium Permanganate, $NaMnO_4$, is prepared in the same way as the potassium salt. It is deliquescent and crystallizes with difficulty. In an impure form it is used as a disinfectant ("Condy's liquid").

DETECTION AND DETERMINATION OF MANGANESE.

Compounds of manganese yield an amethyst borax bead in the oxidizing flame, and a green mass of manganate when fused with sodium carbonate and saltpetre. Manganese compounds are readily detected by these reactions and by the formation of the flesh colored-sulphide.

For determination, manganese is usually precipitated as carbonate, hydroxide or hydrated peroxide, and weighed as Mn_3O_4.

The value of a manganese ore is determined by the amount of chlorine which it evolves from hydrochloric acid.

COBALT. $Co^{II \text{ and } IV}=58.6$.

Occurs native in minute quantities in most meteorites ; combined chiefly with arsenic and sulphur in ores which usually contain also nickel and iron.

Prepared from its oxide by reduction with carbon or hydrogen, or by ignition of its oxalate.

Properties. White metal, harder than iron, malleable, ductile and magnetic. Specific gravity 8.6. Melts at a white heat. In a finely divided state oxidizes in the air, often with ignition. The compact metal is unaltered in air at ordinary temperatures, but oxidizes when heated, and burns at a white heat with a red flame. Decomposes water at a red heat. Readily soluble in nitric acid ; dissolves slowly in dilute hydrochloric or sulphuric acid.

COMPOUNDS OF COBALT.

Cobalt forms two series of compounds of which the cobaltous compounds (CO^{II}) are the more important. They are stable and isomorphous with the corresponding ferrous compounds. In the anhydrous state the salts are violet or blue ; hydrated, of a rose-red tint.

COBALTOUS COMPOUNDS.

Cobaltous Oxide, CoO, is obtained as a greenish powder by heating the hydroxide or carbonate in the absence of air, or the higher oxide to redness in CO_2.

Cobaltous Hydroxide, $Co(OH)_2$, is obtained from a hot cobaltous solution on adding caustic alkali, as a reddish precipitate, which on exposure to the air soon becomes brown. Heated in the air, it is converted into Co_3O_4.

Cobalt Chloride, $CoCl_2$, is formed by heating cobalt in chlorine. Blue crystalline scales which dissolve slowly in water forming a violet solution which becomes reddish on dilution. Cobalt chloride crystallizes as $CoCl_2+6H_2O$ in red prisms which melt on heating to a blue liquid, and are converted on further heating into the blue anhydrous salt. A concentrated solution becomes blue on addition of concentrated hydrochloric acid. Both the anhydrous and crystalline salt are hygroscopic. On account of the change of color due to the loss and absorption of water, the salt has long been used as a sympathetic ink.

Cobalt Sulphate, $CoSO_4+7H_2O$, crystallizes in red prisms readily soluble in water. Forms double salts with alkaline sulphates.

Cobalt Nitrate, $Co2NO_3+6H_2O$. Red prisms, deliquescent and very soluble. Melts below 100° to a violet liquid which at a higher temperature becomes thick and green and then decomposes leaving a residue of Co_3O_4. Used as a reagent in blowpipe analysis and for sympathetic ink.

Cobaltous Sulphide, CoS, is precipitated from a cobalt solution by an alkaline sulphide, as a black powder, nearly insoluble in dilute acids.

Smalt is a potash glass colored intensely blue by cobalt oxide. Formerly much used as a blue pigment, but now largely supplanted by artificial ultramarine.

COBALTIC COMPOUNDS.

Cobaltic Oxide, Co_2O_3. Dark brown powder, formed by gently igniting the nitrate. Converted at a red heat into Co_3O_4; dissolves in acids, yielding brown solutions of very unstable cobaltic salts.

Cobaltic Hydroxide, $Co_2(OH)_6$, is precipitated as a brownish-black powder on adding an alkaline hypochlorite to a cobalt solution.

Cobalto-Cobaltic Oxide, Co_3O_4, is obtained by heating one of the other oxides. It is a non-magnetic, black powder.

Cobaltic Salts. The simple cobaltic salts are very unstable and known only in brown-colored solutions. A number of more stable double salts are known, the most important of which is

Potassium Cobalt Nitrite (Fischer's salt), $K_6Co_2(NO_2)_{12}+3H_2O$. It is obtained as a yellow precipitate when potassium nitrite is added to a cobalt solution acidified with acetic acid :—$2CoCl_2+10KNO_2+4HNO_2=4KCl+2H_2O+2NO+K_6Co_2(NO_2)_{12}$. This reaction is characteristic of cobalt and serves for separation of cobalt from nickel. Fischer's salt or "cobalt yellow" is used as a pigment.

Cobaltamines. A large number of peculiar ammoniacal cobalt compounds are known which are called *cobaltamines.* They may be divided into three chief classes :—

1. *Cobalto-amines.* These consist of a cobaltous salt combined with a number (usually six) of molecules of NH_3.

2. *Cobalti-amines,* cobaltic salts with six, eight, ten or twelve molecules of NH_3.

3. *Oxy-Cobaltamines,* basic cobalt salts with NH_3.

But few of the first and third classes are known, while the members of the second class are very numerous. The constitution of these compounds is as yet imperfectly understood. They are all characterized by peculiar colors and their names are derived from this property.

DETECTION AND DETERMINATION OF COBALT.

Cobalt compounds impart an intense blue color to the borax bead. In the course of analysis, cobalt is always obtained with nickel. (For the separation of cobalt and nickel, cf. nickel). It is usually weighed as Co_3O_4.

NICKEL. $Ni^{II}=58.6$.

Occurs native only in meteorites; combined chiefly with arsenic and with sulphur, almost always accompanied by cobalt.

Extraction. The treatment of nickel ores varies according to the nature of the ores and the other metals which they contain. The extraction and separation from cobalt is always complicated and tedious and usually effected by a wet method. The nickel is finally obtained in the form of oxide from which the metal is reduced by carbon.

Properties. Very hard, white metal; magnetic, malleable, ductile and can be welded. Specific gravity 8.9. Unchanged in the air; decomposes steam slowly at a red heat; dissolves slowly in hydrochloric and sulphuric acids; readily soluble in dilute nitric acid; rendered passive (cf. iron) by concentrated nitric acid. Used chiefly in alloys and for plating metals, usually by electro-deposition.

COMPOUNDS OF NICKEL.

The salts of nickel are usually yellow in the anhydrous state ; in the hydrated state, green.

Nickel Monoxide, NiO, occurs as bunsenite, and may be obtained as a green crystalline powder by heating the hydroxide, carbonate or nitrate. Reduced by carbon or hydrogen at a red heat.

Nickel Hydroxide, $Ni(OH)_2$ is obtained from a nickel solution on adding a caustic alkali as an apple-green precipitate, which dissolves in ammonia with a blue color.

Nickel Peroxide, Ni_2O_3, is obtained as a black powder by gentle ignition of the carbonate or nitrate. The corresponding hydroxide, $Ni_2(OH)_6$, is formed as a black precipitate when a nickel solution is warmed with an alkaline hypochlorite. Both oxide and hydroxide with sulphuric and nitric acids, yield oxygen ; with hydrochloric acid, chlorine ; with ammonia, nitrogen.

Nickel Chloride, $NiCl_2+6H_2O$. Green crystals which on heating become yellow and anhydrous. The anhydrous chloride is also obtained by heating nickel in chlorine, in yellow scales which are difficultly soluble but absorb moisture from the air and then dissolve easily.

Nickel Sulphate, $NiSO_4+7H_2O$. Green prisms, which become $NiSO_4+H_2O$ at 100° ; anhydrous above 280°, and decompose at a red heat. Forms double salts with sulphates of the alkalies of which *ammonium nickel sulphate*, $(NH_4)_2SO_4 + NiSO_4 + 6H_2O$, is employed in nickel plating.

Nickel Nitrate, $Ni2NO_3+6H_2O$. Green deliquescent crystals; decomposes on heating into Ni_2O_3 and finally into NiO.

Nickel Sulphide, NiS. Occurs as millerite and is precipitated from a nickel solution by an alkaline sulphide as a black powder nearly insoluble in dilute acids.

Nickel Cyanide, $Ni(CN)_2$. Obtained as an apple-green precipitate on adding potassium cyanide to a nickel solution. Soluble in excess of potassium cyanide and reprecipitated by dilute acids, even after boiling. (Distinction from cobalt.)

DETECTION AND DETERMINATION OF NICKEL.

Nickel is always obtained in the course of analysis with cobalt and is separated from the latter by the precipitation of cobalt as Fischer's salt; or of the nickel as hydroxide, $Ni_2(OH)_6$, by adding chlorine or sodium hypochlorite to an alkaline solution of the cyanides in excess of potassium cyanide. Nickel is weighed as monoxide.

METALS OF THE CHROMIUM GROUP.

CHROMIUM. MOLYBDENUM. TUNGSTEN. URANIUM.

The metals of this group yield acid-forming trioxides which give very characteristic salts. Chromium is the connecting link with the iron group, its lower oxide closely resembling that of iron.

CHROMIUM. $Cr^{II \text{ and } IV}=52.5$.

Occurs only in combination, chiefly as oxide in chrome iron-stone (FeO,Cr_2O_3); more rarely as lead chromate. It is found in minute quantities in many iron ores and forms the green or red coloring principle of many minerals.

Prepared from its oxide by reduction with carbon; from its chloride by fusion with zinc; or from a solution of chromous chloride by electrolysis.

Properties. Chromium is a grey, crystalline powder. Specific gravity 6.5. Hard as corundum, more difficulty fusible than platinum. Is slowly oxidized when heated in air or steam; burns in the oxy-hydrogen flame with emission of sparks; readily converted into potassium chromate on fusion with potassium chlorate or nitrate. Soluble in hydrochloric and sulphuric acids. Insoluble in nitric acid. The presence of 0.5–0.75 per cent. of chromium in steel renders it very hard.

COMPOUNDS OF CHROMIUM.

Chromium forms three series of compounds: Chromous compounds in which it acts as dyad; chromic compounds in which it acts as tetrad in a hexad double atom (Cr_2^{VI}); and chromic acid compounds. All chromium compounds are brightly colored, hence the name *chromium.*

The *chromous* compounds are very unstable, readily absorbing oxygen and hence acting as powerful reducing agents. They may be obtained by reduction of the chromic compounds.

CHROMIC COMPOUNDS.

Chromic Oxide, Cr_2O_3, occurs as chrome ochre and in chrome iron-stone. Artificially prepared it forms a green powder or crystals. It is infusible; reduced by carbon at a high heat; after ignition almost completely insoluble in acids; dissolved by fused silicates and borates, coloring them green. Used as a green pigment in painting (chrome-green) and in the preparation of colored glass, enamels and porcelain.

Chromic Hydroxide, $Cr_2(OH)_6$, is obtained on adding ammonia to a chromic solution as a pale blue hydrated precipitate which is converted into the oxide at a red heat. It is a weak base and acts as an acid towards strong bases, dissolving in caustic alkalies (cf. aluminium hydroxide), with formation of compounds which are decomposed on boiling.

An hydroxide, $Cr_2O(OH)_4$, is largely used as a pigment (Guignet's green).

Chromic Chloride, Cr_2Cl_6, is obtained by heating an intimate mixture of chromic oxide and carbon in chlorine. It forms peach blossom colored scales which may be sublimed in an atmosphere of chlorine. Heated in the air it is converted into the oxide. Almost insoluble in water, but readily soluble when a trace of chromous, cuprous or stannous chloride is present, yielding a green solution, which is also obtained by dissolving the hydroxide in hydrochloric acid. From this solution readily soluble green crystals, $Cr_2Cl_6+12H_2O$, are deposited on evaporation, which are converted into the anhydrous salt when heated in chlorine or HCl.

Chromic Sulphate, $Cr_2 3SO_4$, forms a violet and a green salt containing different amounts of water of crystallization. The anhydrous sulphate is insoluble in water, and dissolves with difficulty in acids. Chromic sulphate unites with the alkaline sulphates to form chrome-alums of which the most important is

Potassium Chrome-Alum, $Cr_2 3SO_4, K_2SO_4 + 24H_2O$. Crystallizes in large dark purple-red octahedrons which become anhydrous at 350°. The solution heated to 70° becomes dark-green. Used as a mordant and in tanning.

CHROMIC ACID COMPOUNDS.

Chromic Anhydride, CrO_3, is prepared by adding concentrated sulphuric acid to a cold saturated solution of potassium bichromate. The anhydride crystallizes out on cooling in red needles which are very soluble in water, and melt at about 190° to a dark red liquid, which solidifies to a crystalline mass of metallic appearance, and is decomposed at 250° into Cr_2O_3 and oxygen. Chromic anhydride is a very powerful oxidizing agent. With sulphuric acid it yields oxygen; with hydrochloric acid, chlorine. Used in the laboratory as an oxidizing agent.

Chromic Acid and Chromates. *Chromic acid,* H_2CrO_4, is known only in solution, but forms stable salts.

Potassium Chromate, K_2CrO_4, is obtained by adding caustic potash to a solution of the bichromate. Yellow rhombic crystals, readily soluble in water, insoluble in alcohol. On heating, the salt melts without decomposition. The yellow solution becomes red on the addition of acids, from the formation of bichromate. Used as reagent in the laboratory.

Potassium Bichromate, $K_2Cr_2O_7$. By heating chrome iron-stone with potash, normal potassium chromate is formed. The mass is lixiviated with hot water, and the bichromate formed and mostly precipitated by the addition of sulphuric acid. The salt is purified by recrystallization. It forms large red crystals which melt below a red heat and decompose at a white heat into K_2CrO_4, Cr_2O_3 and oxygen. It is much less soluble in water than the normal salt; is reduced in solution by sulphur dioxide to chromic salts and by hydrogen sulphide to Cr_2O_3. Heated with concentrated sulphuric or hydrochloric acid, the solid salt yields chrome alum and oxygen, or chromic chloride and chlorine.

Potassium bichromate is the source of all other chromium compounds and is used as a mordant and oxidizing agent (especially in solution with sulphuric acid), in tanning, in photography and for many other purposes.

Lead Chromate, $PbCrO_4$, occurs naturally and is obtained as a yellow precipitate on adding potassium bichromate to a lead salt. It is largely used as a pigment (chrome yellow), and is frequently employed as an oxidizing agent in organic analysis.

DETECTION AND DETERMINATION OF CHROMIUM.

Chromium compounds yield an emerald green borax bead. On fusion with soda and saltpetre a yellow mass is obtained which is soluble in water; on adding a lead salt to this solution acidified with acetic acid, chrome yellow is precipitated. All chromates are converted by reducing agents into the chromic salts. These reactions, and the color of its compounds and solutions, serve for the detection of chromium. For determination it is precipitated as hydroxide by ammonia from chromic solutions and weighed as chromic oxide.

MOLYBDENUM. Mo=96.

Occurs as sulphide (MoS_2); more rarely as oxide (MoO_3), and as lead molybdate ($PbMo_4$).

Prepared by heating its oxides or chlorides in a current of hydrogen.

White metal, harder than silver. Specific gravity 8.6. Almost infusible. Unaltered in air at ordinary temperatures, oxidized at a red heat to the trioxide. Unattacked by dilute hydrochloric or sulphuric acid; dissolved by nitric acid, concentrated sulphuric acid or aqua regia.

COMPOUNDS OF MOLYBDENUM.

Oxides. Molybdenum forms four oxides: MoO, Mo_2O_3, MoO_2, and MoO_3; and "blue oxides" of various composition.

Molybdenum Trioxide, MoO_3, occurs as molybdic ochre, and may be prepared by roasting the native sulphide, or by heating ammonium molybdate. White powder, which turns yellow on heating; melts at a red heat and volatilizes at a high temperature. Slightly soluble in water; dissolves easily in caustic alkalies and ammonia, yielding *molybdates.* Fused with alkali hydroxides or carbonates, salts are produced which are partly derived from the normal molybdic acid, H_2MoO_4, and partly from polymolybdic acids (e. g. $H_2Mo_2O_7$, $H_2Mo_3O_{10}$).

Molybdic Acid, H_2MoO_4, separates occasionally from a solution of ammonium molybdate containing nitric acid, as a yellow crystalline crust. Soluble in water and acids. On adding zinc to a hydrochloric acid solution, the solution becomes blue, then green and finally dark-brown.

Ammonium Molybdate, $(NH_4)_2MoO_4$, is formed by dissolving the trioxide in ammonia and is precipitated by the addition of alcohol. On evaporation of the solution of the trioxide in ammonia, the *ordinary molybdate of ammonia,* $(NH_4)_6Mo_7O_{24}+4H_2O$, crystallizes out. On heating, both salts give MoO_3.

On adding an excess of a nitric acid solution of ammonium molybdate to a solution containing phosphoric acid, a yellow precipitate is formed, $11MoO_3$, $(NH_4)_3PO_4+6H_2O$, which is insoluble in water and dilute acids; easily soluble in ammonia. This reaction serves for the detection and quantitative separation of phosphoric acid. A similar salt is formed in solutions of arsenic acid.

Molybdenum Chlorides. On heating molybdenum in a current of chlorine, *molybdenum pentachloride,* $MoCl_5$, sublimes in glistening black crystals, which melt at 194° and boil at 268°. The chloride fumes and decomposes in moist air, is deliquescent and easily soluble in water. Heated to 250° in hydrogen, it yields the trichloride, $MoCl_3$, a reddish-brown powder which is insoluble in water, and gives, when strongly heated in carbon dioxide, the dichloride, $MoCl_2$, and the tetrachloride, $MoCl_4$. A number of oxychlorides are also known.

Molybdenum Sulphides. *Molybdenum Disulphide,* MoS_2, occurs as molybdenite which resembles graphite, and is formed as a black powder on heating molybdenum trioxide with sulphur. *Molybdenum trisulphide,* MoS_3, is thrown down as a brown precipitate from acidulated molybdenum solutions by hydrogen sulphide. It forms soluble sulpho-molybdates with the alkalies.

TUNGSTEN. W=183.6.

Occurs in the tungstates of calcium, lead, iron and manganese, the most important natural compound being wolfram, a tungstate of iron and manganese.

Prepared from the trioxide by reduction with carbon or hydrogen; or from the chloride by reduction with sodium vapor. An almost infusible powder. Specific gravity 19.1. Burns at a red heat to WO_3. Heated with acids, it is slowly converted into WO_3. A small quantity of tungsten in steel improves its quality, hardness and tenacity.

COMPOUNDS OF TUNGSTEN.

Oxides. *Tungsten Trioxide,* WO_3, is obtained as a yellow powder by heating tungstic acid Insoluble in water and acids; dissolves readily in caustic alkalies and ammonia with the formation of tungstates. WO_3 is reduced by hydrogen at a red heat to WO_2.

Tungstic Acid, $(HO)_4WO$, is obtained as a white precipitate on adding dilute acids to cold solutions of tungstates; H_2WO_4 when an excess of hot acid is used. On heating it yields, first $H_2W_2O_7$, and at 100°, WO_3. On adding hydrochloric acid and zinc to a solution of a tungstate, the latter is colored blue from the formation of lower oxides.

The tungstates are analogous to the molybdates and are derived partly from the normal acid, H_2WO_4, and partly from poly-acids. The most important salt is *sodium para-tungstate,* $Na_{10}W_{12}O_{41}$, which is prepared by fusion of pulverized wolfram with sodium carbonate, lixiviation of the mass, and crystallization. It forms large colorless prisms containing 28 mol. of water. It is sometimes used in place of sodium stannate as a mordant in dyeing and calico-printing, and is also used for rendering cloth uninflammable.

Tungsten Chlorides. By heating tungsten in chlorine the hexachloride, WCl_6, is obtained as a dark violet crystalline mass, which melts at 275° and boils at 346° ; is decomposed with warm water into tungstic and hydrochloric acids ; heated in hydrogen it yields the chlorides WCl_5, WCl_4 and WCl_2.

Tungsten sulphides resemble those of molybdenum.

URANIUM. $U^{IV \text{ and } VI}=240$.

Occurs chiefly as oxide in pitch-blende.

Prepared from uranous chloride by reduction with sodium. Steel grey, malleable and hard metal. Specific gravity 18.7. Melts at a strong red heat. Burns, on heating, to the green oxide, U_3O_8.

COMPOUNDS OF URANIUM.

Uranium forms two series of compounds : the very unstable uranous compounds, in which it acts as a tetrad ; and the uranic compounds, in which it acts as an hexad. In many compounds the group UO_2 (uranyl) appears in the role of a dyad metal.

URANOUS COMPOUNDS.

Uranium Dioxide, UO_2, is formed when the other oxides are heated in a current of hydrogen. Black powder, which, when heated in the air, is converted into U_3O_8. Soluble in strong acids with the formation of uranous salts.

Uranous Chloride, UCl_4, is obtained by heating uranium in chlorine, or the dioxide in HCl. Sublimes in dark green metallic octahedrons. Very hygroscopic and dissolves in water with hissing to a green solution, which is decomposed on heating.

URANIC COMPOUNDS.

Uranic Oxides. *Uranic oxide,* UO_3, obtained by heating uranic nitrate to 250° as a yellow powder, which dissolves in acids forming uranyl salts.

The Green Oxide, U_3O_8, occurs as pitch-blende, and is formed by heating the other oxides in the air as a green powder. Difficultly soluble in hydrochloric and sulphuric acids; readily soluble in nitric acid.

On strong ignition, the other oxides or ammonium uranate yield the black oxide, U_2O_5, which is used for painting on porcelain.

Uranic Nitrate, UO_22NO_3, obtained by dissolving uranic oxides or pitch-blende in nitric acid. Crystallizes with six mol. of water in large greenish-yellow prisms which are very soluble in water.

Uranyl Ammonium Phosphate, $UO_2NH_4PO_4$, is obtained as a greenish-yellow precipitate, insoluble in acetic acid, by adding a soluble phosphate to a solution of uranyl acetate containing NH_4Cl. This reaction is employed for the volumetric determination of both uranium and phosphoric acid.

Uranates, insoluble in water, are obtained by precipitating a uranyl salt with an alkali. *Sodium Uranate,* $Na_2Ur_2O_7$, is known as uranium yellow, and used for painting on glass and porcelain and for making uranium glass. *Ammonium Uranate* occurs sometimes in commerce and is also called uranium yellow. Heated to redness, it yields U_3O_8, and hence serves as the source of other uranium compounds.

METALS OF THE TIN GROUP.

TIN. TITANIUM. ZIRCONIUM. THORIUM.

These elements are closely connected with the non-metallic element silicon, forming dioxides corresponding to SiO_2, volatile tetrachlorides, and double fluorides corresponding to the silico-fluorides.

TIN. $Sn^{II \text{ and } IV}$ (Stannum)=118.

Occurs only in combination, chiefly as SnO_2 (tinstone); rarely as sulphide with other metals in tin pyrites.

Prepared from tin stone by reduction with coal, and purified from other metals by " liquation."

Properties. White metal of crystalline texture; somewhat harder than lead and very malleable; ductile at 100°, but unfitted for wire on account of its slight tenacity; at 200° very brittle. Melts at 235° and volatilizes at a white heat. Tarnishes very slightly at ordinary temperatures, but becomes covered with a gray suboxide when fused. At a bright red heat it burns with an intense white light to tin oxide. It dissolves in hot hydrochloric acid; in hot concentrated sulphuric acid with evolution of sulphur dioxide, and in cold dilute nitric acid. Is oxidized by strong nitric acid into insoluble metastannic acid.

Used for making stills, kettles and other utensils; as tin foil for a variety of purposes ; for plating iron and copper ; the preparation of important alloys ; as amalgam for silvering mirrors, and as a source of tin salts.

COMPOUNDS OF TIN.

Tin forms two series of compounds : stannous, in which it acts as a dyad, and stannic compounds, in which it acts as a tetrad.

STANNOUS COMPOUNDS.

The stannous salts are obtained by dissolving tin in acids. They possess an unpleasant metallic taste, redden litmus, are readily oxidized to stannic salts, and hence serve as powerful reducing agents.

Stannous Oxide, SnO, obtained by heating stannous oxalate out of contact with air ; or by boiling stannous hydroxide with caustic soda. Dark brown powder which readily takes fire when heated in the air, burning to SnO_2.

Stannous Hydroxide, $Sn(OH)_2$, is precipitated from stannous chloride by sodium carbonate, as a white precipitate soluble in caustic alkalies, but not in ammonia.

Stannous Chloride, $SnCl_2$, obtained by dissolving tin in hydrochloric acid. Crystallizes as $SnCl_2 + 2H_2O$ (tin salt). Carefully heated at 100° the salt becomes anhydrous ; melts at 250°, and at a red heat distils. The anhydrous salt is also obtained by heating tin in a current of HCl. Readily soluble in water, but decomposes when diluted with much water with precipitation of a basic chloride. Forms easily soluble double salts with alkali chlorides. Is a powerful reducing agent. Tin salt is largely used in dyeing and calico-printing and as a reagent in the laboratory.

Stannous Sulphide, SnS, is obtained by heating tin and sulphur as a lead-grey crystalline mass; or as a brown hydrated precipitate on adding hydrogen sulphide to a stannous solution. Soluble in hot hydrochloric acid, insoluble in alkali sulphides, but dissolves on addition of sulphur, or in the poly-sulphides.

STANNIC COMPOUNDS.

Stannic Oxide, SnO_2, occurs as tin-stone, and is formed artificially as a white amorphous powder by heating tin in the air. It is infusible and insoluble in acids. Fused with caustic alkali, it yields alkali stannate. Amorphous stannic oxide is used as a polishing powder and for making milk glass, enamel, and opaque glazes.

Stannic Hydroxide. Two modifications of this compound are known, which probably have the same ultimate composition, but differ in their behavior toward reagents and in the salts which they form. They both act as acids. Both are converted by heat into SnO_2.

Stannic Acid, H_2SnO_3, is precipitated by ammonia from a solution of stannic chloride, or by an acid from a solution of a stannate. It is slightly soluble in water and reddens litmus; soluble in acids, and in alkalies yielding *stannates.*

Metastannic Acid, $H_{10}Sn_5O_{15}=5H_2SnO_3$(?), is obtained by treating tin with concentrated nitric acid as a white powder, insoluble in acids. Soluble in alkalies, but less readily than stannic acid, forming *metastannates,* which possess a complicated composition and crystallize with difficulty.

Sodium Stannate, Na_2SnO_3, is prepared on a large scale by fusing finely divided tin-stone with caustic soda or tin, or with caustic soda and Chili saltpetre. It crystallizes as $Na_2SnO_3+3H_2O$ and is less soluble in hot than in cold water. Used extensively in calico-printing ("preparing salts").

Stannic Chloride, $SnCl_4$, is obtained by heating tin or stannous chloride in a current of chlorine. A colorless, thin, fuming liquid, specific gravity 2.23, which does not solidify at —29° and boils at 120°. Absorbs moisture greedily, solidifying to a soft crystalline mass (butter of tin), $SnCl_4+3H_2O$, which is readily soluble in water. It can be crystallized from its solution with three, five or eight mol. of water of crystallization. On boiling the aqueous solution, HCl is evolved and metastannic acid precipitated. Forms double salts with certain metallic chlorides.

Stannic chloride is employed as a valuable mordant, usually in the form of $SnCl_4+5H_2O$ (oxymuriate of tin). *Pink salt,* $SnCl_4+2NH_4Cl$, and physic, or nitro-muriate of tin obtained by dissolving tin in aqua regia, were formerly much used in dyeing.

Stannic Sulphide, SnS_2, is prepared by subliming a mixture of tin amalgam, sulphur, and sal-ammoniac, in gold-colored scales; and is obtained as a yellow amorphous powder by precipitating a stannic solution with hydrogen sulphide. The precipitate is readily soluble in alkali sulphides forming alkali sulpho-stannates, and in hydrochloric acid. The crystalline sulphide is known as mosaic gold, and used for gilding.

DETECTION AND DETERMINATION OF TIN.

Tin is recognized by its behavior with concentrated nitric acid, and in solution by its sulphides and their reactions. It is readily reduced from its solutions by zinc. Determined as oxide.

TITANIUM. $Ti^{II \text{ and } IV}=50.25$.

Occurs chiefly as oxide (TiO_2) in rutile, brookite and anatase; also in titanates of iron (titanic iron) and calcium, and in minute quantities in many iron ores and other minerals. It is prepared from potassium titano-fluoride by the action of sodium vapor in an atmosphere of hydrogen.

Titanium is a dark grey powder, which assumes, under the burnisher, the lustre and color of iron. Very difficultly fusible; burns brightly, when heated in air, to TiO_2; decomposes boiling water; soluble in dilute acids. Titanium unites directly with nitrogen at a high temperature, forming various *nitrides*.

Titanium Compounds resemble those of tin. *Titanium Chloride,* $TiCl_4$, is prepared by heating a mixture of TiO_2 and carbon in chlorine.

ZIRCONIUM. $Zr^{IV}=90.4$.

Occurs chiefly as silicate in zircon ($ZrSiO_4$). The metal may be obtained as an iron-grey powder by heating potassium zircono-fluoride with potassium; crystallized by heating the same compound with aluminium. The amorphous metal burns readily, when heated, to ZrO_2. In the crystalline state it can only be ignited at the heat of the oxyhydrogen flame. Both modifications are soluble in hydrofluoric acid and oxidized by aqua regia.

Zirconium Chloride, $ZrCl_4$, is prepared in the same way as titanium chloride. It is a white volatile salt. Its solution yields a crystalline basic salt on evaporation.

THORIUM. $Th^{IV}=231.96$.

Occurs chiefly as silicate in a few rare minerals. It is prepared by heating the chloride with sodium, and forms a grey powder, which burns when heated in the air to ThO_2. Readily soluble in nitric acid, but dissolves with difficulty in hydrochloric acid.

Thorium Chloride, $ThCl_4$, is prepared in the same way as the chlorides of titanium and zirconium, and is a slightly volatile deliquescent mass.

VANADIUM. NIOBIUM. TANTALUM.

These three rare metals yield acid-forming pentoxides and other compounds which closely resemble those of the Nitrogen Group; but form no compounds with hydrogen, and are themselves well characterized metals.

VANADIUM. $V^{III \text{ and } V}=51.1$

Occurs as vanadates of lead, copper, etc., in a few rare minerals, and is found in minute quantities in many clays and in certain iron ores. Prepared by reducing the dichloride in a current of hydrogen. It is obtained pure with the greatest difficulty on account of its strong affinity for oxygen at a red-heat. Light grey crystalline powder. Difficultly fusible. Burns at a red-heat to Vd_2O_5. Insoluble in hydrochloric acid; dissolves slowly in hot concentrated sulphuric and in hydrofluoric acids; readily in nitric acid, yielding a blue solution. Fused with caustic soda, sodium vanadate is formed. Heated in nitrogen it is converted into VN.

Vanadium Compounds. Vanadium forms five *oxides* analogous to the oxides of nitrogen: V_2O, V_2O_2, V_2O_3, V_2O_4 and V_2O_5. The first three act as basic oxides; the last two are acid-forming oxides. V_2O_2' enters into many compounds as a triad radicle, and is called *vanadyl* (VO).

Normal Vanadic Acid, H_3VO_4, is not known. Metavanadic acid, HVO_3, forms a fine yellow pigment and is employed in place of gold bronze. Salts of both acids, and also of pyro-vanadic and more complicated poly-acids are known. The meta-vanadates are the most stable of the simpler salts.

Vanadium forms three *chlorides,* VCl_4, V_2Cl_6, V_2Cl_4 and four oxychlorides, $VOCl_3$, $VOCl_2$, $VOCl$, V_2O_2Cl. Vanadium compounds are somewhat used in dyeing and cotton-printing for the production of aniline black, fuchsine, etc.

NIOBIUM. Nb=93.7. TANTALUM. Ta=182.

Niobium and tantalum usually occur together in tantalite, columbite and a few other rare minerals. Neither of them has as yet been prepared in the pure state. Obtained by reduction of potassium fluotantalate with potassium, and of niobium chloride with hydrogen, they form powders which burn when gently heated in air to Ta_2O_5 and Nb_2O_5; in chlorine to chlorides. Both metals are insoluble in aqua regia; niobium dissolves in concentrated sulphuric acid; tantalum only in hydrofluoric acid.

The Chlorides, $NbCl_5$ and $TaCl_5$, are obtained by heating a mixture of the pentoxides and carbon in a current of chlorine. They are yellow, crystalline and volatile substances (melt at 194° and 211°; boil at 240° and 242°) which are decomposed by water into hydrochloric, and niobic ($HNbO_3$) or tantalic ($HTaO_3$) acids.

METALS OF THE GOLD GROUP.

GOLD. PLATINUM. PALLADIUM. RHODIUM.
IRIDIUM. RUTHENIUM. OSMIUM.

Gold and platinum do not combine directly with oxygen; palladium, rhodium and iridium on heating in air or oxygen form oxides, which are decomposed by strong ignition into metal and oxygen; ruthenium and osmium form volatile oxides which do not decompose at the highest temperatures. The lower oxides of this group act as weak bases; the higher ones are acid-forming oxides.

GOLD. $Au^{I \text{ and } III}$ (Aurum)=196.2.

Occurs widely distributed. Usually native in small particles disseminated through quartz rock or the sand produced by the disintegration of the rock; sometimes in considerable masses termed nuggets. Gold is also found combined with bismuth, tellurium, and in minute quantities in many pyrites, blendes and other ores.

Extraction. Native gold is rarely pure, but almost always contains silver, occasionally in considerable amount, and usually traces of iron and copper, and other metals of the gold group.

1. Native gold is usually separated from the rock or sand in which it occurs, (a) by washing away the latter with water. The heavy gold alloy remains and is often amalgamated with mercury in order to complete the separation (panning and hydraulic mining), the mercury being afterwards removed by volatilization. (b) Gold is sometimes extracted from roasted aurifrous pyrites by converting it into gold chloride, by the action of chlorine, dissolving the chloride in water and precipitating the gold by green vitriol. (c) From ore in which the gold is chemically combined the metal is extracted by alloying it with lead, which is then removed by cupellation.

2. The gold is purified from silver, (a) by dissolving the latter in nitric acid. In order to separate the two metals by this method the alloy must not contain more than 35 per cent. of gold; if richer than this, the necessary quantity of silver must be added. The silver is recovered from the silver nitrate solution by precipitation with copper. (b) More usually by means of the cheaper sulphuric acid. In this case the alloy must not contain more than 25 per cent. of gold. The granulated alloy is heated with concentrated sulphuric acid until the silver (and copper) is completely dissolved. The sulphur dioxide produced in this operation is often used for the preparation of sulphuric acid. The silver is precipitated from the silver sulphate solution by copper (or by iron).

The gold still retains traces of other metals which are removed by passing chlorine into the molten metal; silver chloride is formed and rises to the surface, while the chlorides of other metals are volatilized. Chemically pure gold may be prepared from the aqua regia solution of the metal by precipitation with a reducing agent.

Properties. Yellow metal, crystallizing in cubes or octahedrons. Softer than silver. Specific gravity 19.3. The most malleable and ductile of all metals. Melts at 1100° forming a bluish-green liquid and contracts considerably on solidification. Volatile at the highest temperatures. Finely divided gold, precipitated from solution by reducing agents, is of a dull brown color. When precipitated in a dilute solution in so fine a state that it remains suspended, it appears purple red or blue. Is unattacked at any temperature by oxygen, water or fused chlorate of potassium; does not dissolve in any single acid with the exception of selenic acid. Fused nitrates or alkaline hydroxides act upon it, and it unites readily with chlorine, and is hence dissolved by solutions which evolve chlorine (aqua regia, etc). Bromine acts like chlorine; but iodine does not sensibly attack it. Sulphur and hydrogen sulphide do not act upon it; but gold sulphide is readily formed by the action of fused alkali poly-sulphides.

Uses. Gold is used for gilding, usually by electro-deposition from a solution of potassium auro-cyanide, $KAu(CN)_2$. For most purposes gold is alloyed with silver or copper or with both. The fineness of the alloy is often expressed in *carats*, twenty-four carat gold being pure.

COMPOUNDS OF GOLD.

Gold forms two series of compounds, aurous and auric, in which it acts as monad and triad respectively.

AUROUS COMPOUNDS.

Gold Monoxide, Au_2O, is precipitated as a dark violet powder by the action of caustic potash on aurous chloride. It is decomposed at 250° into gold and oxygen. Insoluble in sulphuric and nitric acids, readily dissolved by aqua regia.

Aurous Chloride, $AuCl$, is formed by heating auric chloride to 180°. Decomposed at a higher heat into gold and chlorine. Insoluble in water, but decomposed by boiling water into gold and auric chloride.

AURIC COMPOUNDS.

Gold Trioxide, Au_2O_3, is obtained by drying the hydroxide as a blackish brown powder which is readily decomposed on heating into gold and oxygen.

Gold Hydroxide, $Au(OH)_3$, is obtained by heating a solution of auric chloride with an excess of magnesia and washing the precipitate with nitric acid. Brown powder; decomposes gradually in the light and quickly on heating. Dissolves in caustic alkalies forming unstable salts, called aurates, derived from the meta-acid, $AuO.OH$. Ammonia converts it into a greenish or brownish powder, which when dry, explodes violently on heating or percussion (fulminating gold). This compound is also formed by precipitating auric chloride with ammonia or ammonium carbonate.

Auric Chloride, $AuCl_3$. Moist chlorine at ordinary temperatures, or dry chlorine at 300° converts finely divided gold into $AuCl_3$. (Chlorine unites with gold to form $AuCl_2$, which is decomposed at 300° or by water into $AuCl$ and $AuCl_3$). The chloride is also formed by dissolving gold in aqua regia. Dark brown crystalline mass. Readily soluble in water, alcohol and ether. Its aqueous solution is reduced by a current of hydrogen and somewhat by light. With hydrochloric acid and soluble chlorides it forms soluble double salts. Auric chloride is the most important compound of gold. It is used for the preparation of gilding solutions; as a source of gold preparations and also in medicine.

Sodium Gold Chloride, $AuCl_3,NaCl + 2H_2O$, is used in photography and in porcelain and glass painting.

Gold Disulphide, Au_2S_2, is obtained as a hydrated precipitate by leading hydrogen sulphide into a cold solution of the chloride. Insoluble in nitric or hydrochloric acid; soluble in aqua regia and dissolves in alkali sulphides with the formation of sulphosalts.

Purple of Cassius is precipitated on adding a mixture of stannous and stannic chlorides to an auric chloride solution, as a brown, purple or dark violet powder of varying composition, which is used in staining glass and in porcelain painting,

DETECTION AND DETERMINATION OF GOLD.

Gold is readily detected by the formation of the purple of Cassius, or by reduction from its solutions. It is always weighed in the metallic state.

PLATINUM. $Pt^{\text{II and IV}}=194.2$

Platinum occurs only native, but almost invariably alloyed with the other metals of this group and with iron, copper, etc., chiefly in river sand and detritus.

Extraction. The native platinum alloy, after separation from the sand or earth in which it occurs by washing, is treated with cold dilute aqua regia which dissolves the gold; and then heated with concentrated aqua regia, usually under pressure, which dissolves the platinum, palladium, rhodium and a little iridium. This solution is then evaporated and the residue heated to 125° to convert the palladium and rhodium salts into lower chlorides. The mass is then dissolved in water and the platinum (with a small amount of iridium) precipitated as platinum ammonium chloride; this precipitate is dried and converted by heating into spongy platinum, which is pressed and finally fused in a lime crucible by the oxyhydrogen blowpipe.

Properties and Uses. Tin-white metal, soft as copper, less tenacious than iron, highly malleable and ductile. Can be welded at a white heat. Specific gravity 21.5. Fusible only at the temperature of the oxyhydrogen flame. Is not attacked at any temperature by oxygen, water or any single acid. Fused alkaline hydroxides, sulphides and cyanides, and saltpetre attack it. Soluble in aqua regia.

Platinum is extensively used for chemical utensils. For this purpose it is often alloyed with iridium which renders it harder and less easily attacked by chemical agents. *Platinum Sponge* obtained by heating platinum ammonium chloride, and *platinum black* obtained by reducing a solution of platinum with zinc, etc., have the property of condensing gases, especially oxygen, and thus act as strong oxidizing agents.

COMPOUNDS OF PLATINUM.

Platinum Oxides. *Platinum Monoxide*, PtO, obtained by carefully heating the hydroxide, $Pt(OH)_2$ is a black powder, insoluble in dilute acids, decomposed at a strong heat into platinum and oxygen.

Platinous Hydroxide, $Pt(OH)_2$, obtained by decomposing the dichloride with hot caustic potash. Dark brown powder.

Platinum Dioxide, PtO_2, obtained by gently heating *platinic hydroxide* $Pt(OH)_4$, which is prepared by adding caustic potash to a platinic chloride solution and washing the precipitate with acetic acid. The hydroxide dissolves easily in acids (except acetic) forming platinic salts. Towards alkalies it acts as an acid, forming *platinates*.

Platinum Chlorides. *Platinous Chloride*, $PtCl_2$, obtained by heating platinic chloride to 300°, is a greenish-grey powder, insoluble in water, decomposed into its elements by heat.

Platinic Chloride. By dissolving platinum in aqua regia, and evaporation, a brownish-red crystalline mass, $PtCl_4,HCl+6H_2O$, is obtained, which dissolves readily in water, alcohol and ether. On adding silver nitrate to a cold aqueous solution, $PtCl_4,2AgCl$ is precipitated, which is decomposed by hot water into $AgCl$ and a solution of $PtCl_4$, Double chlorides with the alkalies and alkaline earths are readily formed, most of which are difficultly soluble in water and insoluble in alcohol. The sodium double salt is however soluble in alcohol, water and ether.

Platinic chloride also forms many characteristic double salts with organic compounds.

Used as an important reagent in the laboratory, in photography, for platinizing metals, porcelain or glass ; for making platinum sponge, platinum black and other platinum preparations.

Platinum Sulphides. *Platinum Monosulphide,* PtS, is formed by heating platinum sponge and sulphur. Insoluble in acids and aqua regia, but dissolves in alkali sulphides. On heating, it is decomposed into its elements. *Platinum Disulphide,* PtS_2, is obtained as a black precipitate on leading hydrogen sulphide into platinic chloride solution. Converted into chloride by aqua regia, into sulphate by fuming nitric acid and dissolves in alkali sulphides. Heated out of contact with air it gives PtS; in air, platinum.

DETECTION AND DETERMINATION OF PLATINUM.

Platinum is readily reduced from its compounds and identified by its behavior. For quantitative determination it is precipitated as platinum ammonium chloride and weighed as platinum.

PALLADIUM. $Pd^{\text{II and IV}} = 106.2$

Occurs in most platinum ores and is also found with gold. Prepared by the ignition of palladium cyanide, which is precipitated from the aqua regia solution by mercury cyanide.

Properties. Resembles platinum, but is softer, lighter and more fusible. Volatilizes in the oxyhydrogen flame with a green vapor. On heating it tarnishes at first, but becomes bright again at a higher temperature. Soluble in nitric acid. Absorbs large quantities of hydrogen when heated to redness in this gas, or brought in contact with nascent hydrogen, forming palladium hydride, Pd_2H_4, which acts as a powerful reducing agent. At a high heat the hydrogen is expelled. Palladium is used to some extent as a substitute for silver, and by dentists as a substitute for gold. Palladium compounds closely resemble those of platinum.

RHODIUM. $Rh^{\text{II and IV}} = 104.1$

Occurs with platinum. Prepared from the solution from which the platinum has been precipitated by ammonium chloride. It resembles aluminium in color, is less fusible than platinum, non-volatile, malleable. Specific gravity 12.1. Insoluble in all acids. Dissolves in aqua regia only when alloyed with platinum or other metals. Heated to a high temperature in the air it oxidizes superficially. It forms three oxides, RhO, Rh_2O_3, RhO_2, and salts which are derived from Rh_2O_3.

IRIDIUM. $Ir^{\text{II and IV}} = 192.5$.

Occurs with platinum and is obtained from the residue insoluble in aqua regia. White, brittle metal; at a white heat somewhat malleable. Specific gravity 22.42. Much less fusible than platinum. Insoluble in acids. Dissolves in aqua regia only in the form of iridium black, or when alloyed with much platinum. It is converted into oxide by heating with potassium hydroxide, chlorate or nitrate. Used for making platinum iridium alloys, and an alloy with osmium. Used for making pen points and in the form of oxide for a black porcelain color. The iridium compounds are very similar to those of platinum. It forms also a sesquioxide and sesquichloride, Ir_2O_3 and Ir_2Cl_6.

RUTHENIUM. Ru=103.5.

Occurs with platinum and is obtained from the residue insoluble in aqua regia. It is hard and brittle ; less fusible than iridium. Specific gravity 12.26. Oxidizes slowly in the air. Insoluble in acids and aqua regia. Oxidized by fused potassium hydroxide, chlorate or nitrate. Ruthenium forms a number of oxides, and two series of salts.

OSMIUM. Os=195.

Occurs with the other platinum metals and is prepared with them. It is easily separated from the other metals of the platinum group by the formation of a very volatile tetroxide, OsO_4. Specific gravity 22.48. Infusible. In a finely divided state it oxidizes readily when heated in the air to OsO_4, which is also produced by nitric acid or aqua regia. It is used only in an alloy with iridium for tipping gold pens.

Osmium Tetroxide or "osmic acid," OsO_4, is obtained by the oxidation of the metal. Sublimes in glistening needles at a moderate heat. The fused oxide boils at 100°, yielding a colorless vapor which has a penetrating odor, and which attacks the eyes and lungs most violently. Soluble in water. Reducing and organic substances precipitate black osmium from this solution, which is used on this account for blackening microscopical preparations.

APPENDIX.

SPECIFIC GRAVITY OF THE METALS.

Osmium,	22.48	Indium,	7.42
Iridium,	22.42	Tin,	7.29
Platinum,	21.50	Niobium,	7.2
Gold,	19.32	Zinc,	7.15
Tungsten,	19.1	Chromium,	6.8
Uranium,	18.7	Antimony,	6.71
Mercury,	13.55	Cerium,	6 68
Ruthenium,	12.26	Didymium,	6.54
Rhodium,	12.1	Lanthanum,	5.1
Thallium,	11.85	Gallium,	5.95
Lead,	11.37	Vanadium,	5.5
Palladium,	11.4	Zirconium,	4.15
Thorium,	11	Barium,	3.75
Silver,	10.57	Aluminium,	2.6
Tantalum,	10.4	Strontium,	2.54
Bismuth,	9.8	Beryllium,	2.07
Copper,	8.92	Caesium,	1.88
Nickel,	8.9	Magnesium,	1.74
Cadmium,	8.6	Calcium,	1.57
Cobalt,	8.6	Rubidium,	1.52
Molybdenum,	8.6	Sodium,	.98
Manganese,	8.0	Potassium,	.87
Iron,	7.86	Lithium,	.59

MELTING POINTS AND BOILING POINTS OF THE METALS.

	Melts at		Melts at	Boils at
Osmium,	2500	Lanthanum,		
Iridium,	2500	Uranium,		
Molybdenum,		Strontium,		
Chromium,		Calcium,		
Rhodium,	2000	Magnesium,	750	1100
Platinum,	2000	Aluminium,	600	
Tungsten,		Antimony,	440	1090–1450
Manganese,	1900	Zinc,	433	1040
Ruthenium,	1800(?)	Lead,	335	1450–1600
Palladium,	1700	Cadmium,	315	860
Iron (pure),	1600	Thallium,	288	red heat.
Cobalt,	1500	Bismuth,	270	1090–1450
Nickel,	1450	Tin,	230	1450–1600
Barium,		Lithium,	180	
Zirconium,		Indium,	176	red heat.
Gold,	1100	Sodium,	95.6	861–954
Copper,	1100	Potassium,	62.5	719–731
Silver,	1040	Rubidium,	38.5	
Beryllium,		Gallium,	30.15	
Didymium,		Caesium.	26.5	
Cerium,		Mercury,	—39.5	357

ORDER OF TENACITY, MALLEABILITY AND DUCTILITY OF THE COMMONER METALS.

Tenacity.	*Malleability.*	*Ductility.*
Iron.	Gold.	Gold.
Copper.	Silver.	Silver.
Platinum.	Aluminium.	Platinum.
Silver.	Copper.	Aluminium.
Gold.	Tin.	Iron.
Zinc.	Platinum.	Nickel.
Nickel.	Lead.	Copper.
Tin.	Zinc.	Zinc.
Lead.	Iron.	Tin.
	Nickel.	Lead.

CONDUCTIVITY OF THE METALS.

For Heat.		*For Electricity.*	
Silver,	100	Silver,	100
Copper,	73.6	Copper,	89.9
Gold,	53.2	Gold,	70.9
Magnesium,	34.3	Aluminium,	49.7
Aluminium,	31.3	Magnesium,	36.3
Zinc,	28.1	Zinc,	27.2
Cadmium,	20.06	Cadmium,	22.5
Tin,	15.2	Tin,	13.3
Iron,	11.9	Iron,	12.6
Lead,	8.5	Platinum,	9.8
Platinum,	8.4	Lead,	7.8
Antimony,	4.03	Antimony,	3.3
Bismuth,	1.8	Mercury,	1.6
Mercury,	1.35	Bismuth,	1.3

ELECTRO-CHEMICAL SERIES.

Negative End.
Oxygen,
Sulphur,
Selenium,
Nitrogen,
Fluorine,
Chlorine,
Bromine,
Iodine,
Phosphorus,
Arsenic,
Chromium,
Vanadium,
Molybdenum,
Tungsten,
Boron,
Carbon,
Antimony,
Tellurium,
Tantalum,
Titanium,
Silicon,
Hydrogen,
Gold,
Osmium,
Iridium,
Platinum,
Rhodium,
Palladium,
Mercury,
Silver,
Copper,
Bismuth,
Tin
Lead,
Cadmium,
Cobalt,
Nickel,
Iron,
Zinc,
Manganese,
Uranium,
Cerium,
Thorium,
Zirconium,
Aluminium,
Didymium,
Lanthanum,
Yttrium,
Beryllium,
Magnesium,
Calcium,
Strontium,
Barium,
Lithium,
Sodium,
Potassium,
Caesium,
Rubidium.
Positive end.

COMPOSITION OF SOME OF THE MORE IMPORTANT ALLOYS.

Coins. *Gold Coins.* United States, France and Germany: Gold 90, copper 10. England: Gold 22, copper 2.

Silver Coins. United States, France (5 franc piece) and Germany: Silver 90, copper 10. France (small coins): Silver 83.5, copper 16.5. England: Silver 92.5, copper 9.5.

Nickel Coins: Copper 75, nickel 25. *Copper Coins:* Copper 95, tin 4, zinc 1.

Brass: Copper 67–72, zinc 33–28. Brass is malleable, ductile, and can be rolled. Next to iron it is the most important "metal."

Pinchbeck: Copper 90, zinc 10.

German-Silver: Copper 50, zinc 25, nickel 25.

Bronzes. *Gun-metal:* Copper 90, tin 10. *Bell-metal:* Copper 78, tin 22. *Speculum-metal:* Copper 67, tin 33. *Aluminium-bronze:* Copper 90–95, aluminium 10–5. Aluminium-bronze has the appearance of gold and unites the advantages of bronze and brass. The alloy containing 5 per cent. aluminium is as hard and elastic as steel, but less brittle, non-magnetic and not liable to rust. *Phosphor-bronze* is a bronze containing 0.25–2.5 per cent. phosphorus, and is especially hard, tough and elastic.

Pewter: Tin 92, lead 8. *Common Solder:* Tin 50, lead 50. *Coarse Solder:* Tin 33, lead 67. *Fine Solder:* Tin 67, lead 33.

Britannia-metal: Tin 90, antimony 10. Usually contains 1 to 3 per cent. zinc and about 1 per cent. copper.

Type-metal: Lead 80, antimony 20. Varies considerably in composition.

Shot: Lead with 0.3–0.8 per cent. arsenic.

Wood's Metal: Lead 4, bismuth 5–8, tin 2, cadmium 1–2. Melts at 66–72°.

Rose's Metal: Bismuth 2, tin 1, lead, 1, Melts at 93°.7.

AMALGAMS.

Amalgams are usually prepared by the direct union of the metal with mercury, with or without the aid of heat. They are also produced when mercury is added to a solution of a metallic salt; or the metal placed in a solution of mercuric nitrate; or by electrolysis of a metallic solution by a weak current with mercury as the negative electrode.

Amalgams are liquid, pasty or hard, according to the proportion of mercury which they contain. They are decomposed by great pressure or heat, though some retain a portion of the mercury at a temperature above the boiling point of mercury.

The more important amalgams are :

Sodium Amalgam ; used in extracting gold and silver, and as a reducing agent in the laboratory.

Tin Amalgam ; used for silvering mirrors.

Cadmium and Copper Amalgams ; used for stopping teeth.

SOLUBILITIES, SPECIFIC GRAVITIES AND MELTING POINTS OF CERTAIN COMPOUNDS.

	100 parts of water dissolve at			
	20°	100°	Spec. Grav.	Melts at
Alum (Ammonium) cryst.,	13.6	422	1.63	
Alum (Potassium) cryst.,	15	375.5	1.72	
Aluminium Sulphate cryst.,	107	1132	1.62	
Ammonium Chloride,	37.3	72.8	1.52	
Ammonium Nitrate,	200		1.7	
Ammonium Sulphate cryst.,	76.3	97.5	1.77	
Barium Chloride,	37.7	58.8	3.85	
Barium Hydroxide,	5	50	4.5	
Barium Nitrate,	9.2	32.2	3.23	
Cadmium Chloride,	140.8	142.9	3.26	541
Calcium Chloride,	74	153	2.22	719
Calcium Oxide,	0.13	0.077	3.08	
Calcium Sulphate,	0.21	0.17	2.32	
Cobalt Sulphate,	36.4	65.7(70°)	3.53	
Copper Sulphate cryst.,	42.3	203.3	2.27	
Ferrous Sulphate cryst.,	115(24°)	370(90°)	1.88	
Lead Chloride,	0.8	5	5.8	498
Lead Nitrate,	56	140	4.41	
Magnesium Chloride cryst.,	130	366	1.56	708
Magnesium Sulphate,	36.2	73.8	2.65	
Manganese Sulphate cryst.,	123	92	2.09	
Mercuric Chloride,	7.4	54	5.42	287

	100 parts of water dissolve at			
	20°	100°	Spec. Grav.	Melts at
Potassium Bichromate,	12.4	94.1	2.70	
Potassium Carbonate,	94	153.7	2.29	834
Potassium Chloride,	35	59	1.98	734
Potassium Chromate,	62.9	79.1	2.72	
Potassium Chlorate,	7.2	55.9	2.33	359
Potassium Hydroxide,	200		2.04	
Potassium Nitrate,	31	247	2.09	339
Potassium Sulphate,	10.9	26.2	2.64	
Potassium Acid Sulphate,		79.1		210
Silver Nitrate,	227	1111	4.34	218
Sodium Borate (Borax) cryst.,	7.9	201.4	1.72	561
Sodium Carbonate,	25.9	48.5	2.47	814
Sodium Acid Carbonate,	9.6			
Sodium Chloride,	36	39.6	2.15	772
Sodium Hydroxide,	133	250	2.13	
Sodium Nitrate,	88	178.2	2.24	316
Sodium Phosphate cryst.,	17.2		1.54	888
Sodium Sulphate,	52.8	42.5	2.65	861
Strontium Hydroxide cryst.,	1.8	100	1.39	
Strontium Nitrate,	70.8	101.1	2.93	645
Zinc Sulphate cryst.,	161.5	653.6	2.01	

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